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ABSTRACT

One of twelve water treatment technology units, this student manual on chemistry/bacteriology provides instructional materials for twelve competencies. (The twelve units are designed for a continuing education training course for public water supply operators.) The competencies focus on the following areas: waterborne diseases, water sampling procedures, manganese test and remedies, iron test and remedies, hardness test and remedies, pH testing methods, coagulation processes, jar test, corrosion control by pH adjustment, alkalinity test, dry chemical feeders, and chemical solution feeders. For each competency, student objectives, resources (information sheets), learning activities, and evaluation activities are provided. The appendices contain information sheets on bacteriological sampling and examination, softening hard water, coagulation, corrosion and control, and laboratory equipment. Generally, an evaluation form with answer key is provided for each student competency. (JH)

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Water Treatment Technology

CHEMISTRY/BACTERIOLOGY

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December, 1978

U.S. DEPARTMENT OF HEALTH,
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CE 022 313

CHEMISTRY/BACTERIOLOGY

EVALUATOR'S COPY

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EVALUATOR'S INSTRUCTIONS

Name ________________________________

Address ________________________________

____

Phone ________________________________

This is your EVALUATOR'S COPY of the CHEMISTRY/BACTERIOLOGY unit. It contains all of the materials received by the operators who will be working on the unit. It also contains *answer keys* and *directions* for the evaluation of each Competency. For each Competency Sheet, there is a COMPETENCY, a list of RESOURCES, LEARNING ACTIVITIES, and EVALUATION ACTIVITIES.

Your student operators will call you when they are ready to be evaluated. An operator may complete the evaluations in *any* order. S/he may choose to complete one at a time, or several at one time. *It is up to you and each operator to set convenient schedules for the evaluations.* It is important that the operators learn the skills listed in each Competency. It is your job to be sure the operators can perform the Competencies.

If an operator does not complete an evaluation correctly, s/he may try again later after practicing the skills. It would be helpful for you to tell the operators just how they should improve their knowledge and skills in order to complete the evaluation successfully. Be sure to call if you have any questions.

This package contains:

- - - - - - CHEMISTRY/BACTERIOLOGY unit

- - - - - - Evaluation Forms for each Competency

You will receive instructions as to where to send the completed evaluation forms. Please be sure that the operator's name and your name are on the forms, as well as the date on which the evaluation was performed.

INSTRUCTIONS FOR THE STUDENT OPERATOR

The CHEMISTRY/BACTERIOLOGY unit contains twelve competencies. A "competency" is a skill that is necessary for good on-the-job performance. Each competency in this unit is necessary for you to perform on the job according to the West Virginia Department of Health. There is a Competency Sheet to help you understand each competency.

DIRECTIONS FOR COMPLETING THE UNIT:

1. Read each Competency Sheet carefully. Every Competency Sheet contains:

   COMPETENCY
   Objectives
   RESOURCES
   LEARNING ACTIVITIES
   EVALUATION ACTIVITY (IES)

2. After reading the whole Competency Sheet, you should understand what will be expected of you.

3. Do the LEARNING ACTIVITIES that you think will help you the most. This may include reading some of the RESOURCES which have been included with the unit.

4. If you need additional help, ask an operator in your plant for assistance, or call your evaluator; he's there to help you.

5. When you are ready to complete the EVALUATION ACTIVITY for one or more competencies --- call your evaluator. The two of you will schedule the best time for your evaluation.

   a. If you complete the evaluation correctly, go on to the next competency.
   b. If you do not complete the evaluation correctly, your evaluator will tell you what your mistakes were. You must go back and practice your skills and read the RESOURCES again. Call your evaluator when you are ready to try again. *You must wait at least one week before being re-evaluated on a competency.*

SEE NEXT PAGE

6. When you complete all your evaluations correctly you will receive a Certificate of Accomplishment and 2.5 Continuing Education Units of Credit.

7. If you want to go on to another unit, contact your evaluator for the materials. The procedure will be similar for all units.

AN EXPLANATION OF CONTINUING EDUCATION UNITS

For student operators in the OPERATOR-IN-TRAINING category who will be taking the certification examination for Class I, this course is an opportunity to increase your knowledge and competence in the treatment of water for human consumption. Successful completion of the manuals will also, it is hoped, make passing the certification exam much easier. The Continuing Education Units which you will get by taking this course will not decrease the minimum one year operating experience requirement for a Class I certificate.

For student operators who now hold a Class I certificate, the completion of all twelve manuals in the course will reduce the currently required three years of total operating experience for Class II certification by approximately four months. Thus, if you complete the course, you will need only two years and eight months of total operating experience before you take the Class II exam.

Here is a listing of the Continuing Education Unit (CEU) value of each of the manuals.

Manual	CEU Value
General Plant Operation	1.4
Taste, Odor, and Color	0.6
Chemistry/Bacteriology	2.5
Chlorination	2.5
Fluoridation	0.6
Filtration	1.2
Hydraulics	0.6
Distribution Systems	1.2
Pumps	0.6
Wells	1.0
Cross Connections	0.4
Springs	0.4
Total	13.0

For comparison with other types of training previously available, the CEU conversions below are given.

Public Water Supply Operator's Short Course (given by State Department of Health Personnel)	4 CEUs
One Semester Hour of College Credit	1.5 CEUs
One Year of College Credits (30 semester hours)	45 CEUs

(Chemistry/Bacteriology #11)
Competency Sheet

COMPETENCY: Identify and describe the cause and effect of water borne diseases which can make people ill.

Objective: Identify water borne diseases by name.

Objective: Describe the causes and effects of various water borne diseases.

Objective: Describe your legal responsibility related to water borne diseases as a water treatment plant operator.

RESOURCES:

(a) "Water - Borne Diseases", Resource Sheet.

(b) Public Water Supply Regulations, West Virginia State Board of Health.

LEARNING ACTIVITIES:

(a) Identify those specific West Virginia regulations which relate to water borne diseases. What are your responsibilities as a water plant operator?

(b) Read the Resources listed above to find information about water borne diseases. Learn to recognize the causes and symptoms of the various water borne diseases.

EVALUATION ACTIVITY: Complete Evaluation Form #11. You will be expected to know your legal responsibility as a West Virginia water treatment plant operator related to water-borne diseases. You must also be able to identify the causes and symptoms of various water borne diseases. You may use references.

WATER BORNE DISEASES

Illnesses from minor ailments to serious epidemics can be transmitted to people by water. Things which cause disease and are transmitted are:

Pathogenic microorganisms- *bacteria, viruses, protozoa*

Parasitic life- *worms*

Metallic poisoning- *arsenic, barium, cadmium, chromium, cyanide, lead, mercury*

Insecticides

Detergents

Radiation

Water supplies are exposed to contamination by pathogenic bacteria. Bacterial diseases spread by water are:

Typhoid fever- transmitted through infected water or milk, etc.
Cause: Excretal pollution of water supplies.
Symptoms: Fever, intestinal disorders.

Cholera- transmitted through infected water, uncooked vegetables, and other foods.
Cause: Fecal contamination from a carrier of the disease.
Symptoms: Violent diarrhea, vomiting, muscular cramps, collapse.

Bacterial dysentery- transmitted through infected water by flies.
Cause: Excretal pollution of water supplies.
Symptoms: Diarrhea, fever.

Tularemia- transmitted through infected water, especially where rodents are present.
Cause: Contamination from infected animals (especially rabbits).
Symptoms: Inflammation of lymph glands, irregular fever, aching.

Pseudomonas aeruginos (re: Tuberculosis)- transmitted through sewage effluents from tuberculosis sanitariums.
Cause: May be danger of infection from bathing, by inhalation of atomized water during a storm, or from irrigation of crops with effluent or sewage from a sanitarium.
Symptoms: Formation of tubercles in various body tissues, especially the lungs; general deterioration.

Amoebic dysentery- transmitted by contaminated food or water.
Cause: Cyst that is excreted in the feces of carriers of the disease. Flies may convey the infection to food.
Symptoms: Inflammation, abdominal pain, toxemia, bloody diarrhea.

Gastroenteritis (including all types of Salmonella organisms)- transmitted through infected water or food.
Cause: Improper chlorination; inadequate protection of water supplies from surface drainage and seepage from sewage; cross-connections; back-siphonage.
Symptoms: Inflammation of the lining of the stomach and intestines.

Water supplies are exposed to contamination from viruses. Viral diseases spread by water are:

Poliomyelitis- transmitted through infected water accompanied by gross pollution.
Cause: Virus entering through mouth.
Symptoms: Infection of the pharynx and the rest of the alimentary tract.

Infective hepatitis- transmitted through infected water usually accompanied by gross fecal pollution. Extremely contagious.
Cause: Virus in water or spread by person-to-person contact. Some infections traced to fecally contaminated wells and drain leaks in spring water supply.
Symptoms: Inflammation of the liver.

Water supplies are exposed to contamination from Heavy Metals. They are:

Arsenic
Barium
Cadmium
Chromium
Cyanide
Lead
Mercury
Selenium
Silver

Barium, cadmium, and mercury occur in nature. They are dissolved by water to form various salts that may pollute water supplies.

Arsenic, chromium and cyanide are contaminants from industry. Lead is found as a result of corrosive water dissolving lead from lead pipes or fixtures. Selenium and Silver are also contaminants from industry.

Danger to food in contact with polluted water- A wide variety of diseases can be transmitted through food which has been contaminated by polluted water. Foods which are eaten raw or only partially cooked are frequently responsible: shellfish, fish, water cress, and other vegetables, fruits.

(Chemistry/Bacteriology #11)
Evaluation Form
EVALUATOR'S COPY
Page 1 of 2

Operator ____________________

Evaluator ____________________

Date ____________________

Directions: In the space provided write the word "True" or "False". If you answer "False" write the correct answer on the lines under the question.

1. *True* Public water supplies in W. Va. should not contain any impurities which may be hazardous to consumers.

2. *True* Any water plant operator who violates any W. Va. water treatment regulation may be fined.

3. *True* Improper chlorination may result in an outbreak of gastroentritis.

4. *False* Infective hepatitis can not be spread by drinking water.
Some infective hepatitis has been traced to the pollution of spring drinking water supply through a drain leak; some traced to fecally contaminated wells.

5. *True* Cholera is not a problem in W. Va. because of protection of water supplies.

6. *True* Diarrhea is a symptom of bacterial dysentery.

(Chemistry/Bacteriology #11)
Evaluation Form
EVALUATOR'S COPY
Page 2 of 2

7. *False* Polluted water used to water milk cows does not effect the quality of milk.

Not only milk but milk products such as cheese and ice cream could spread disease if cattle are watered with sewage-polluted water.

8. *True* Shellfish and water cress could spread disease from sewage polluted water.

9. *True* Fever and headaches are two of the symptoms of Salmonella typhimurium.

10. *True* Tularemia is a water borne disease which causes inflammation of the lymph glands.

11. *False* Tuberculosis can not be spread by drinking water.

Tuberculosis can infect water sources which receive the sewage effluent from a tuberculosis sanitarium.

12. *True* Some water borne diseases may also be spread by insects.

(Chemistry/Bacteriology #12)
Competency Sheet

COMPETENCY: Demonstrate the ability to gather, package, and send representative water samples for performing bacteriological and chemical lab tests.

Objective: State your legal responsibility as a plant operator for submitting water samples for bacteriological and chemical analysis.

Objective: Gather, package, and send representative water samples to the West Virginia State Hygienic Laboratory using proper equipment and procedures.

RESOURCES:

(a) Public Water Supply Regulations, West Virginia State Board of Health.

(b) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter V. *-Blue Appendix-*

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about collecting water samples. Identify those West Virginia regulations that relate to gathering, packaging, and sending water samples for examination. Note all West Virginia procedures and the correct address to send samples to the State Hygienic Laboratory.

(b) Following correct procedures, gather, package, and send water samples to the West Virginia State Hygienic Laboratory for 1) bacteriological analysis and 2) chemical, physical, and radiological analysis. If possible have an experienced operator observe you while you are collecting and packaging water samples. Make sure the samples you collect are representative of your water.

Remember that chemists at the W. Va. State Hygienic Laboratory will answer any questions you may have about the results of your lab tests.

EVALUATION ACTIVITY: You will be expected to (a) state your legal responsibility as a water plant operator for submitting water samples for bacteriological and chemical analysis and (b) demonstrate that you can correctly gather, package, and send representative water samples to the W. Va. State Hygienic Laboratory. You must use the proper equipment and procedures.

(Chemistry/Bacteriology #12)
Evaluation Form
Page 1 of 2

Operator ______________________

Evaluator ______________________

Date ______________________

______ Yes   The operator correctly stated his/her legal responsibility for submitting water samples for bacteriological and chemical analyses.

Answer should be to the effect that: samples should be collected from representative areas of each public water supply distribution system and submitted to the W. Va. State Hygienic Laboratory in standard sterilized water containers ***provided by the laboratory****. (If the treatment plant has its own* ***certified*** *laboratory, submission of samples may be waived by written permission from the W. Va. Division of Sanitary Engineering.)*

Population Served	*Minimum Number of Samples Per Month**
25 to 999	*1*
1,000 to 2,500	*2*
2,501 to 4,000	*4*
4,001 to 6,000	*7*
6,001 to 8,000	*9*
8,001 to 10,000	*11*
10,001 to 25,000	*30*
25,001 to 50,000	*60*
50,001 to 70,000	*75*
70,001 to 100,000	*100*
100,001 to 130,000	*110*
130,001 to 160,000	*120*
160,001 to 190,000	*130*
190,001 to 220,000	*140*
220,001 to 250,000	*150*
250,001 to 290,000	*160*
290,001 to 320,000	*170*

**The State Director of Health may require that additional samples be submitted.*

______ Yes   The operator gathered, packaged, and sent representative water samples to the W. Va. State Hygienic Laboratory for bacteriological and chemical analyses. (If the plant has its own certified lab omit the "sending" from the demonstration above.)

--Sampling stations should be selected which give representative results of water in the area.
--Outside spigots are recommended. Chrome fixtures (usually inside) may be damaged by properly flaming the fixture to sterilize it.
--Samples should ***not*** *be collected from fire plugs.*
(a) Let the water run for several minutes, then shut it off.

(b) Flame the faucet to destroy any bacteria that may be clinging to it. The flame should be directed against the inside edge of the faucet. Do not overheat the fixture.
(c) Open the faucet. Let the water run in a stream about the size of a pencil.
(d) While the water is running, remove the cap from the collection bottle. Do not touch the top or inside of the bottle. Do not lay the top down.
(e) Hold the bottle under the faucet. Try not to splash much water on the lip of the bottle.
(f) Remove the bottle from under the faucet and replace the cap.
(g) Fill out the identification completely. Replace the bottle in the shipping case.
(h) Ship the package to the W. Va. State Hygienic Laboratory in South Charleston, W.Va. The samples should be collected as near to shipping time as possible.

(Chemistry/Bacteriology # 13)
Competency Sheet
Class II only *

COMPETENCY: Determine the amount of manganese present in your raw and finished water.*

Objective: State the maximum allowable concentration of manganese (Mn) in finished water in West Virginia.

Objective: Identify problems caused by excessive concentrations of manganese in finished water.

Objective: Perform the procedures necessary to determine the amount of manganese present in your raw and finished water.

Objective: Describe steps which can be taken to lower an excessive concentration of manganese.

RESOURCES:

(a) W. Va. Public Water Supply Regulations, adopted by the West Virginia State Board of Health.

(b) "Iron and Manganese Removal", Resource Sheet I, Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual No. 85-13, June 15, 1968.

(c) "Manganese Test", Resource Sheet II.

(d) Laboratory Equipment and Supply Requirements For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES:

(a) Use your W. Va. Public Water Supply Regulations to determine the *maximum allowable concentration of manganese in your finished water*.

(b) Using the Resources listed above, describe symptoms that would lead you to suspect excessive manganese concentrations. What problems does excessive manganese cause?

(c) Using the "Manganese Test", Resource Sheet and any other help you need, perform tests to determine the amount of manganese present in your raw and finished water.

Check to see if the manganese concentration in your finished water is below the maximum allowable concentration according to the W. Va. Public Water Supply Regulations.

**This competency is required for Class II operators and persons applying for Class II certification. Not required for Class I.*

(d) Describe the steps that are necessary to lower excessive concentrations of manganese. If the concentration of manganese is too high what specifically would you do to lower it? Which procedure would be best to use?

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #13. You will be expected to answer questions about allowable concentrations of manganese, how to identify symptoms of excessive manganese, and how to lower excessive manganese concentrations. You may not use references.

(b) Demonstrate for your evaluator your ability to determine the amount of manganese present in your raw and finished water. Check to make sure that the manganese level in the finished water is below the maximum allowable concentration. You may use references.

IRON AND MANGANESE REMOVAL*

Iron

(1) General. The presence of iron in excess of 0.3 ppm causes unpleasant taste and reddish color on plumbing fixtures. If in excess of 2.0 ppm, it is very objectionable. There are three general classes of iron-bearing waters: those that begin to precipitate iron immediately after aeration; those in which iron is not easily precipitated by aeration; and those which precipitate part of the iron readily but hold the remainder indefinitely. High iron content is frequently encountered in ground water supplies; surface waters from swampy areas may also contain considerable iron. If the amount of ferrous iron does not exceed 2 to 3 ppm, color and fixture staining can be prevented by treatment with a polyphosphate; above this range, removal is necessary. About 2 ppm of polyphosphate are required to sequester 1 ppm of iron. Iron present in water in soluble ferrous form is removed by oxidizing it to insoluble ferric hydrate. Methods of removal are influenced by presence of organic matter, manganese, carbon dioxide, and carbonates. Methods used include aeration, contact beds, filters, special zeolites, and chemical treatment. Iron removal plants generally consist of aerator units, small settling basins, and sand filters.

(2) Aeration. In waters high in carbon dioxide and containing no dissolved oxygen, iron exists as ferrous bicarbonate. When the water is aerated, carbon dioxide is released, iron changes to ferrous hydroxide and, on contact with oxygen, is oxidized to ferric hydrate. Ferric hydrate is an insoluble flocculent form of iron, readily removed by sedimentation and filtration. The oxygen required is 0.14 ppm per ppm of iron oxidized. Aeration to oxygen saturation is not recommended with low alkalinity water because it increases corrosiveness. Spray, wood, and coke-tray type aerators are efficient and require little attention in operation. Aeration for tastes and odors also will often oxidize iron. Air-lift pumps, if used in wells, may provide all the aeration required.

(3) Contact Aeration. High organic-matter content may necessitate use of contact aerators to remove iron. The action consists of accumulation of the organic matter on a coarse medium such as stone or coke, with accompanying oxidation of the iron which deposits on the medium.

*Source: Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual #85-13, June, 1968, pp. 168-169, 175.

(4) Chemical Treatment:

(a) Dissolved iron is effectively removed by lime or lime-soda treatment for softening water. The high pH created by lime results in formation of a rapid-settling ferric hydrate floc. Oxidizing chemicals such as chlorine or hypochlorites can be used to supply the necessary oxidation to convert iron to the insoluble form, although aeration prior to chemical addition usually is much more efficient and economical.

(b) Where iron is easily oxidized by aeration the following chemical additions hasten subsequent sedimentation (best dosage is determined by laboratory flocculation tests):

1. Lime and aluminum sulfate or ferric salts with 20 to 30 minute flocculation. Recarbonation should follow if the lime treatment has increased the pH sufficiently to throw water out of chemical balance.
2. Aluminum sulfate.
3. A pH over 8.0 helps flocculate the iron hydrate and may, in case of high iron content, produce all the floc necessary to remove turbidity. Lime is used for this pH adjustment. This treatment is necessary in acid waters.

(5) Sand Filters. Where iron resists oxidation by aeration and sedimentation, sand filters act as contact beds and remove most of the remaining iron. The oxidizing action is confined to the upper portion of the bed and results in coating the sand with iron oxide, which has been found to be more active than clean sand for iron removal. Precautions in operation of filters should be observed.

Manganese

Manganese, when present, is almost invariable associated with iron. Due to its chemical similarity to iron, it is removed by the same processes. However, manganese is not oxidized as readily as iron. This is shown by black manganese stains below the iron stains in sand filters.

(1) Filter sand with a manganese dioxide coating aids in removing the manganese. Contact aerators are quite effective.

(2) Since manganese is removed by coagulation only at a relatively high pH value, it is necessary to use a coagulant which has its optimum in a high range. Ferrous sulfate is usually used in a pH range above 9.0.

(3) Oxidation of the manganese with chlorine, followed by coagulation and pH control, sedimentation, and filtration, usually gives satisfactory results.

(4) Polyphosphates will hold manganese in solution provided the manganese does not exceed approximately 1 ppm. A weight ratio of about 2 ppm to 1 ppm of manganese is required.

IRON AND MANGANESE REMOVAL BY ION EXCHANGE

a. Zeolite Softening. Cation or base exchange zeolites will remove iron and manganese within the desirable limits of 0.3 ppm. Thus a sodium zeolite will also remove these two minerals in the process of softening and, is some instances, the softening process becomes secondary to the principal objective of iron and manganese removal. All iron and manganese must be in solution in the ferrous or manganous state before passing through a zeolite bed of this type; otherwise a precipitate will form that will clog the bed. Therefore, air lift pumps or other aerating devices cannot be used ahead of the zeolite beds.

b. Manganese Zeolite Process. This method is used to remove iron and manganese that remain in small amounts (less than 0.1 ppm) after other processes have removed the bulk of these two minerals. The procedure is not one of base exchange but of contact oxidation in which a zeolite treated with manganous sulphate and potassium permanganate oxidizes the minerals from the soluble state to the insoluble state, depositing the oxides formed on the grains. Regeneration is with potassium permanganate, which restores the oxidizing capacity of the bed. This method is normally limited to waters used for special processing and industrial uses.

MANGANESE TEST*

Purpose

Manganese causes laundry and staining problems similar to iron. Manganese stains and deposits are often darker and harder to get rid of. Manganese may be a problem at levels as low as 0.1 mg/l.

Equipment

---10 or more 250 ml erlenmeyer flasks.
---A 25 ml buret (or appropriate pipets) for measuring the standard manganese solution.
---A 100 ml graduated cylinder for measuring distilled water and sample.
---A pipet for measuring the mixed reagent solution.
---A gas burner or electric hot plate.
---One or more wire gauze squares, 20 mesh.
---Glass beads.
---One or more glass stirring rods.
---10 or more 100 ml nessler tubes, matched, tall form; and support.
---Wash water bottle for rinsing flask, beads, and stirring rod.

Reagents

A. Stock Manganese Solution

1) Weigh 0.308 gram dry manganous sulfate monohydrate ($MnSO_4 \cdot H_2O$). Transfer the weighed chemical to a 250 ml beaker. Dissolve in 100 ml distilled water.
2) Carefully add one ml concentrated sulfuric acid (H_2SO_4) with constant mixing.
3) Transfer the acid solution to a one liter volumetric flask. Rinse the beaker with three 100 ml portions of distilled water. Dilute further to the one liter mark with distilled water. Stopper and mix thoroughly.

B. Standard Manganese Solution*

1) Use a volumetric pipet to measure 10 ml stock manganese solution into a 100 ml volumetric flask.
2) Dilute to the 100 ml mark with distilled water.
3) Stopper and mix thoroughly.

**Prepare the standard manganese solution on the day you will use it.*

*Source: Adapted from Simplified Procedures for Water Examination, American Water Works Association, AWWA No. M12, 1975, pp. 78-80.

C. Mixed Reagent Solution

1) Measure 200 ml distilled water and pour into a 1,500 ml beaker.
2) Measure 400 ml concentrated nitric acid.
3) While stirring with one hand, slowly and *carefully* add the nitric acid to the distilled water.
4) Weigh 75 grams mercuric sulfate ($HgSO_4$) and dissolve in the acid solution prepared in Step 3.
5) Meaure 200 ml 85% phosphoric acid (H_3PO_4) and mix into solution #4 above .
6) Weigh 0.035 gram silver nitrate ($AgNO_3$), add to solution (#5 above) and mix to dissolve.
7) Allow the solution to cool to room temperature. Then transfer it to a one liter graduated cylinder. Dilute to the 1,000 mark with distilled water. Mix thoroughly by pouring back into the beaker and stirring.

D. Ammonium Persulfate Crystals or Powder $(NH_4)_2S_2O_8$

Procedure

The sample must be free of turbidity or color which cannot be eliminated by preliminary boiling.

A. Preparation of Manganese Standards

1) Prepare the following series of manganese standards. Measure the indicated volumes of standard manganese solution (B above) into separate 250 ml flasks:

Standard Manganese Solution ml	Manganese mg/1
0	0
0.5	0.05
1.0	0.1
2.0	0.2
3.0	0.3
4.0	0.4
6.0	0.6
8.0	0.8
10.0	1.0

2) Use a graduated cylinder to measure 75 ml distilled water and add to each flask.

B. Mix the sample thoroughly. Measure 100 ml into a 250 ml flask which has been marked with a line at the 90 ml level.
C. Add 5 ml mixed reagent solution to the sample.

D. Heat the sample to boiling. Boil until the liquid level reaches the 90 ml mark. *Prevent spattering and bumping of the boiling liquid by placing a wire gauze over the heat source. Rest the bottom of the flask on the wire gauze. As an added safeguard, add 4 or 5 glass beads to the flask or place a glass stirring rod in the flask.*

E. Add 5 ml mixed reagent solution to each manganese standard prepared in step A above. Mix.

F. Weigh 1 gram portions of ammonium persulfate crystals. Add these portions to each manganese standard and to the sample.

G. Heat the contents of each flask to boiling; boil gently for one more minute. *A pink-to-purple color will develop in the manganese standards at this stage.*

H. Remove each flask from the heat. Allow to stand for one minute.

I. Cool each flask to room temperature under running tap water. Do not delay. Slow cooling may cause some loss in permanganate color.

J. Transfer the contents of each flask to a separate 100 ml nessler tube. Rinse the flask, glass beads, and stirring rod with 5-10 ml distilled water. Add the wash to the nessler tube.

K. Add enough distilled water to bring the volume to the 100 ml mark.

L. Mix each nessler tube by turning it upside down 4-6 times.

M. Compare the sample with the standards. Determine the mg/l manganese present from the pink-to-purple color.

(Chemistry/Bacteriology #13)
Evaluation Form
EVALUATOR'S COPY
Part I
Page 1 of 2

Operator ______________________

Evaluator ______________________

Date ______________________

1. The maximum allowable concentration of manganese for finished water leaving my treatment plant is *0.05* mg/l.

The following questions are to be answered "True" or "False" in the space provided. If you answer "False" write the correct answer on the lines below the question.

2. *False* Too much manganese causes the water to look very dark blue.
Excessive manganese causes a reddish/brown color.

3. *True* Usually if there is too much manganese in the water there is also too much iron in the water.

4. *True* Manganese in water can cause stains on clothes and plumbing fixtures.

5. *False* The best way to lower an excessive manganese concentration is to chlorinate heavily.
Heavy chlorination may cause manganese deposits in all types of pipelines and seriously increase friction head losses.

6. *True* A high pH level is necessary to remove manganese by coagulation.

7. *False* Manganese is more easily oxidized than iron.

COMPETENCY: Determine the amount of iron present in your raw and finished water.

Objective: State the maximum allowable concentration of iron (Fe) in finished water in West Virginia.

Objective: Identify problems caused by excessive concentrations of iron in finished water.

Objective: Perform the procedures necessary to determine the amount of iron present in your raw and finished water.

Objective: Describe steps which can be taken to lower an excessive concentration of iron.

RESOURCES:

(a) W. Va. Public Water Supply Regulations, adopted by the West Virginia State Board of Health.

(b) "Iron and Manganese Removal", Resource Sheet I from Competency #13.

(c) "Iron Test", Resource Sheet.

(d) Laboratory Equipment and Supply Requirements For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES:

(a) Use your W. Va. Public Water Supply Regulations to determine the *maximum allowable concentration of* iron *in your finished water.*

(b) Using the Resources listed above, describe symptoms that would lead you to suspect excessive iron concentrations. What problems does excessive iron cause?

(c) Using the "Iron Test" Resource Sheet and any other help you need, perform tests to determine the amount of iron present in your raw and finished water.

Check to see if the iron concentration in your finished water is below the maximum allowable concentration according to the W. Va. Public Water Supply Regulations.

(d) Describe the steps that are necessary to lower excessive concentrations of iron. If the concentration of iron is too high what specifically would you do to lower it? Which procedure would be best to use?

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #14. You will be expected to answer questions about allowable concentrations of iron, how to identify "symptoms" of excessive iron concentrations, and how to lower excessive iron concentrations. You may not use references.

(b) Demonstrate for your evaluator your ability to determine the amount of iron present in your raw and finished water. Check to make sure that the iron level in the finished water is below the maximum allowable concentration. You may use references.

IRON TEST*

Purpose

Iron should be removed from water where it causes color, staining, bacterial, and taste and odor problems. This method is designed to measure the total iron that is normally present in water or that may be introduced as a result of treatment with iron coagulants or by corrosion of iron pipe.

---A glass sample container with a glass or plastic top is preferable to one with a metal cover.
---Thorough shaking of the bottle is important in order to suspend and distribute all of the iron throughout the sample.
---In order to prevent contamination, a flask, tube, beaker or any other glassware which has held chemicals with a high iron content should be cleansed by boiling for at least 1/2 hour with a solution made of equal volumes of concentrated hydrochloric acid and distilled water.
---The following substances can cause *errors* in this test:
   a) zinc concentrations greater than 10 times the iron value
   b) copper and cobalt greater than 5 mg/l
   c) nickel more than 2 mg/l
   d) bismuth, silver, cadmium, mercury, molybdate, and large amounts of strong oxidizing agents.

Equipment

---10 or more 100 ml nessler tubes, matched, tall or low form; and support.
---A 25 ml buret or appropriate pipets for measuring the standard iron solution.
---A 100 ml graduated cylinder for measuring the sample. (A 50 ml or 25 ml graduated cylinder may be used for smaller samples with particulate matter which may clog a pipet.)
---One or more 250 ml erlenmeyer flasks.
---A safety or automatic pipet for measuring concentrated hydrochloric acid.
---Pipets for measuring hydroxylamine hydrochloride, acetate buffer, and phenanthroline solutions.
---A gas burner or electric hot plate.
---One or more wire gauze squares, 20 mesh.
---Four or more glass beads.
---One or more glass stirring rods.
---A wash water bottle for rinsing flask, beads, and stirring rod.

Reagents

A. Stock Iron Solution

---

*Source: Adapted from Simplified Procedures for Water Examination, Amer. Water Works Assoc., AWWA No. M12, 1975, pp. 64-68.

1) Use a 50 ml graduated cylinder to measure 50 ml distilled water and pour into a 250 ml beaker.
2) Use the same 50 ml graduated cylinder to measure 20 ml concentrated sulfuric acid.
3) While stirring with one hand slowly and *carefully* add the sulfuric acid to the distilled water. Heat is generated by mixing sulfuric acid and distilled water. Pour slowly and mix well to avoid dangerous spattering. Allow the solution to cool to room temperature.
4) Weigh 0.7022 grams dry ferrous ammonium sulfate $Fe(NH_4)_2(SO_4)_2 \cdot 6H_2O$ and add to the solution from #3 above. Dissolve the crystals by stirring the solution.
5) Weigh 1 gram potassium permanganate ($KMnO_4$). Place in a 250 ml beaker. Dissolve in 100 ml distilled water.
6) Add solution (#5) dropwise to solution (#4). Stir constantly until a faint pink color persists.
7) Transfer solution (#6) to a one liter volumetric flask. Rinse the beaker with three 100 ml portions of distilled water. Dilute further to the one liter mark with distilled water. Stopper and mix thoroughly.

B. Standard Iron Solution*

1) Use volumetric pipet to measure 10 ml stock iron solution into a 100 ml volumetric flask.
2) Dilute to the 100 ml mark with distilled water. Stopper and mix thoroughly.

**Prepare the standard iron solution on the day you will use it because the solution is unstable.*

C. Concentrated Hydrochloric Acid, HCl

D. Hydroxylamine Hydrochloride Solution

1) Weigh 10 grams hydroxylamine hydrochloride ($NH_2OH \cdot HCl$) and place in a 250 ml beaker.
2) Use a graduated cylinder to measure 100 ml distilled water and pour into the 250 ml beaker. Stir to dissolve the crystals.

E. Acetate Buffer Solution

1) Weigh 250 grams ammonium acetate ($NH_4C_2H_3O_2$). Place in a one liter beaker and dissolve in 150 ml distilled water.
2) Add 700 ml glacial (or concentrated) acetic acid. Mix.
3) Transfer the mixed ammonium acetate-acetic acid solution to a one liter graduated cylinder. Dilute to the 1,000 ml mark with distilled water. Mix thoroughly by pouring back into the beaker and stirring.

F. Phenanthroline Solution

1) Weigh 0.1 g of 1,10 phenathroline monohydrate ($C_{12}H_8N_2 \cdot H_2O$) and transfer to a 250 ml beaker.

2) Use a graduated cylinder to measure 100 ml distilled water and add to the beaker.

3) Add 2 drops of concentrated hydrochloric acid to the beaker.

4) Stir the solution to dissolve. *If the crystals fail to dissolve quickly, place the beaker over a hot water bath to hasten dissolution.* Throw away the solution when it darkens on prolonged standing.

Procedure

A. Preparation of Iron Standards

1) Prepare the following series of iron standards. Measure the indicated volumes of standard iron solution into separate 100 ml nessler tubes:

Standard Iron Solution ml	Iron mg/l
0	0
0.5	0.05
1.0	0.10
2.0	0.20
3.0	0.30
4.0	0.40
6.0	0.60
8.0	0.80
10.0	1.0

2) Add distilled water to the 100 ml mark, and mix.

B. Mix the sample thoroughly and measure the correct sample volume for the indicated range:

Sample Volume ml	Iron Range mg/l
100	0.05-1.0
50	1.2 -2.0
25	2.1 -4.0

1) Place the sample in a 250 ml flask.

C. Add 2 ml concentrated hydrochloric acid, and mix.

D. Add 1 ml hydroxylamine hydrochloride solution, and mix.

E. Heat the solution to boiling, then boil for 5 minutes. *Prevent spattering of the boiling liquid by putting a wire gauze over the heat source, then resting the bottom of the flask on the wire gauze. For further protection add 4 or 5 glass beads to the flask or place a glass stirring rod in the flask.*

F. Cool the sample to room temperature.
G. Transfer the sample to a 100 ml nessler tube. Rinse the flask, glass beads, and stirring rod with 5-10 ml distilled water. Add this wash water to the nessler tube.
H. Add enough distilled water to bring the volume to the 100 ml mark.
I. Using a measuring pipet, add 10 ml acetate buffer solution to the sample and series of standards prepared earlier. Mix thoroughly by turning each tube upside down 4-6 times.
J. Let the tubes stand at least 15 minutes for the color to develop.
K. Compare the sample with the standards. *Determine the amount of total iron present from the pink color.*
L. Calculate the mg/l total iron by multiplying the result found in step "K" by the correct factor:

Sample Volume ml	Multiply mg/l iron by:
100	1
50	2
25	4

(Chemistry/Bacteriology #14)
Evaluation Form
EVALUATOR'S COPY
Part I

Operator ______________________________

Evaluator ______________________________

Date ______________________________

1. The maximum allowable concentration of iron for finished water leaving my treatment plant is *0.3* mg/l.

The following questions are to be answered "True" or "False" in the space provided. If you answer "False" write the correct answer on the line(s) below the question.

2. *True* Excessive iron causes water to taste bad.

3. *True* Excessive iron causes reddish stains on plumbing fixtures.

4. *False* Ground water supplies rarely have excessive iron concentrations.

   Ground water supplies very frequently have high iron concentrations.

5. *False* Iron cannot be removed with chemical treatment.

   Dissolved iron can be removed by lime or lime-soda treatment for softening water.

(Chemistry/Bacteriology #15)
Competency Sheet
Class II*

COMPETENCY: Determine the amount of "hardness" for your raw and finished water.

Objective: Identify problems caused by excessively "hard" water.

Objective: Perform the procedures necessary to determine the amount of "hardness" for your raw and finished water.

Objective: Identify various procedures for softening water which is too hard.

RESOURCES:

(a) "Hardness Test" Resource Sheet.

(b) "Softening Hard Water" adapted from Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual No. 85-13, June 15, 1968, pp. 167-168. *-Yellow Appendix-*

(c) Laboratory Equipment and Supply Requirements For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES:

(a) Using the information on the Resource Sheet and Appendices, practice performing the procedure for testing the "hardness" of your water. If possible, have an experienced operator observe you while you practice.

(b) Identify "symptoms" of excessively hard water. Why would your customers prefer soft water?

(c) Using the Appendices as references, describe the most appropriate procedure(s) for softening water at your treatment plant (if softening is desirable).

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #15. It contains questions about identifying symptoms of excessive water hardness, causes of hardness, and procedures for treatment. You may use references.

(b) You will be expected to perform the procedures necessary to determine the amount of "hardness" for your raw and finished water while your evaluator observes you. You may use references.

**This competency is required for Class II operators and persons applying for Class II certification. Not required for Class I.*

"HARDNESS TEST"*

Purpose

Calcium and magnesium causes hardness in water. The hardness can be of two types: *carbonate* and *noncarbonate.* Hardness can be removed by treatment with lime, soda ash, cation exchanger, or a combination of all three.

Special Note

This method is designed for routine hardness determinations of potable water.

---The water should be free of color or turbidity which might affect the indicator response.
---The titration should be performed on a sample which is at room temperature. The reaction is slow in a cold sample. The indicator decomposes in heat.
---The titration should be completed within 5 minutes after the buffer is added. This avoids trouble caused by the precipitation of calcium carbonate.
---The buffer and solid indicator mixture are subject to deterioration. They should be kept tightly stoppered when not in use.
---*A poor or green "off-color" endpoint means a need for fresh indicator.*

Equipment

---A 25 ml buret and support.
---A 50 ml graduated cylinder or appropriate volumetric pipets for measuring.
---Two or more 250 ml flasks or porcelain casseroles.
---A dropping pipet (or medicine dropper), 0.5-1 ml, for dispensing the buffer.
---Two small measuring spoons, 0.2-0.3 ml capacity, for dispensing the sodium cyanide crystals and the dry indicator mixture.

Reagents

The following prepared solutions and solid mixtures are available commercially:

A. Sodium Cyanide Crystals, NaCN

Handle this poison carefully with a small spoon or spatula. Do not get in mouth. Avoid inhalation--fumes may be deadly.

*Source: Adapted from Simplified Procedures for Water Examination, American Water Works Association, AWWA No. M12, 1975, 61-64.

B. Buffer Solution

1) Carefully weigh 1.179 grams dry disodium ethylene-diaminetetraacetate dihydrate (abbreviated EDTA or Na_2EDTA) of reagent quality.

Carefully weigh 0.780 grams magnesium sulfate ($MgSO4 \cdot 7H_2O$).

Carefully transfer the two weighed chemicals to a 100 ml beaker and dissolve in 50 ml distilled water.

2) Weigh 16.9 grams ammonium chloride (NH_4Cl) and place in a 400 ml beaker. Use a 250 ml graduated cylinder to measure 143 ml concentrated ammonium hydroxide and add to the 400 ml beaker. Dissolve the ammonium chloride in the concentrated ammonium hydroxide.

3) Pour the solution from step (1) into solution from step (2) while stirring.

4) Pour solution from step (3) back into the 250 ml graduated cylinder and dilute with distilled water to the 250 ml mark. Mix the solution thoroughly by pouring it back and forth between the 400 ml beaker and the 250 ml graduated cylinder. *Store in a tightly stoppered bottle.*

C. Solid Indicator Mixture

1) Weigh 0.5 grams Eriochrome Black T dye. Weigh 100 grams sodium chloride (NaCl). Place both chemicals in a mortar and grind together with a pestle until the dark dye is uniformly distributed throughout the white salt. *Store in a tightly stoppered bottle.*

D. EDTA Titrant

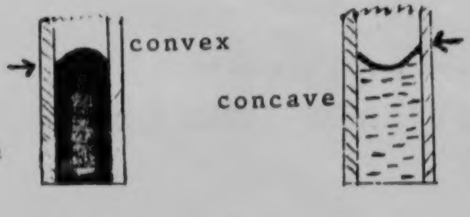
1) Weigh 3.723 grams dry disodium ethylene-diaminetetraacetate dihydrate (abbreviated EDTA or Na_2EDTA) of reagent grade quality. Carefully transfer the solution to a one liter volumetric flask. Rinse the beaker with three 100 ml portions of distilled water. Dilute further with distilled water to the one liter mark. Stopper and mix thoroughly. *EDTA titrant loses strength slowly; the loss is generally negligible during the first six months but becomes noticeable after about a year.*

Procedure

A. Fill the buret with EDTA titrant. Record the liquid level in the buret by reading at the bottom of the meniscus*.

B. Measure the correct sample volume for the indicated hardness ranges:

*Meniscus: the curved upper surface of a column of liquid



Sample Volume ml	Hardness Range mg/1 as $CaCO_3$
50	0 - 300
25	301 - 600
10	601 - 1,500

If a sample of only 25 ml is needed, add 25 ml distilled water to bring the total volume to 50 ml; with a 10 ml sample, add 40 ml distilled water. Measure the additional distilled water with a graduated cylinder. Place the sample (and extra distilled water if needed) in a 250 ml flask (or porcelain casserole).

C. Prepare a color comparison blank: place 50 ml distilled water in a similar 250 ml flask (or porcelain casserole).

D. Measure out 0.25 gram sodium cyanide crystals and add to the color comparison blank and the sample. Dissolve the crystals by mixing.

E. Add 1-2 ml of the buffer solution to the color comparison blank and the sample. Mix thoroughly.

F. Add 1 measure (0.2 g) solid indicator mixture to the color comparison blank and the sample. Mix to dissolve.

G. Add 1 drop of EDTA titrant at a time from the buret to the color comparison blank until the purplish color changes to bright blue. *Sometimes no drops are needed; other times as many as 4 drops may be needed.* Record the new buret level by reading at the bottom of the meniscus.

H. If the sample turns a red or purple color in step (G) above, gradually add EDTA titrant from the buret. Shake the flask constantly (or stir the contents of the porcelain casserole constantly). Continue adding the titrant until the red color turns to a purplish tinge. Stop adding titrant at this point for ten seconds but continue shaking (or stirring).

I. Resume adding the EDTA titrant drop by drop until the purple color turns to the same bright blue in the comparison blank. Shake the flask. The color change from purple to bright blue will occur between 1 to 4 drops.

J. Record the new buret level by reading at the bottom of the meniscus.

K. Calculate the gross volume of titrant used: subtract the initial buret reading (G) from the last reading (J).

L. Calculate the blank correction by subtracting the buret reading in step (A) above from the buret reading in step (G).

M. Calculate the net volume of titrant used for the sample alone by subtracting the result found in step (L) from the result found in step (K).

N. Calculate the hardness in terms of mg/l as calcium carbonate by multiplying the result from step (M) by the correct factor:

Sample Volume ml	Multiply ml EDTA Titrant by:
50	20
25	40
10	100

Note: Water containing over 200 ppm of hardness expressed as $CaCO_3$ is considered very hard; 100-200 ppm is hard; 50-100 ppm is moderately hard; below 50 ppm is *soft*.

(Chemistry/Bacteriology #15)
Evaluation Form
EVALUATOR'S COPY
Part I
Page 1 of 2

Operator ______________________________

Evaluator ______________________________

Date ______________________________

Directions: Write the word "True" or "False" in the space provided next to each item. If you answer "False" write the correct answer on the lines under the question.

1. *True* Calcium and magnesium cause hardness in water.

2. *False* Water which contains about 200 ppm of hardness is very *soft*.
Water containing over 200 ppm of hardness is considered very hard.

3. *True* Hard water causes stains on glassware and kitchen utensils.

4. *True* Sedimentation and filtration are useful procedures in softening water.

5. *True* The purpose of lime-soda process plants is to soften water.

6. *True* Carbonate hardness is often called "temporary hardness".

7. *True* In the "split treatment" process a large portion of raw water is overtreated.

8. *True* Treatment with excess lime and recarbonation can reduce carbonate hardness.

(Chemistry/Bacteriology #16)
Competency Sheet

COMPETENCY: Determine the pH levels for your raw and finished water.

Objective: Describe what the pH level represents.

Objective: Describe how the pH level affects water treatment.

Objective: Perform the procedures necessary to determine the pH level of your raw and finished water.

RESOURCES:

(a) "What is pH?" - Resource Sheet I.

(b) pH Test Resource Sheet II.

(c) pH Test Resource Sheet III.

(d) Laboratory Equipment and Supply Requirements For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES

(a) Read Part I of the Resource Sheet to find a definition of pH value and suggested procedures for testing for the pH level of your water.

(b) Practice performing the procedures necessary to determine the pH levels of your raw and finished water (Resource Sheet Part II). If possible, have an experienced operator observe you while you practice these procedures. Does the pH level differ between the raw water and the finished water? Based on your results of the pH level of your raw water what treatment would you recommend? How would chlorination, coagulation and softening be affected if the pH level of your raw water was used throughout the treatment process?

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #16. It contains questions about the definition of pH value and the purpose of determining the pH levels of your raw and finished water. You may not use references.

(b) You will be expected to correctly perform the procedures necessary to determine the pH levels of your raw and finished water while your evaluator observes you.

WHAT IS pH?

West Virginia Department of Health

Environmental Health Services Laboratory

Division of Sanitary Engineering

pH is a number that exactly describes the degree of acidity or basicity of a solution. We can make a good analogy with the measurement of temperature. Here we have the terms "cold"and "hot" and we immediately realize that these are very general terms which cannot be used with any degree of accuracy. Accordingly, the temperature scale was developed and now a temperature reading of, say 50° C, means the same to everyone and is scientifically accurate.

In the same way, the pH scale was developed. Centuries ago, man discovered that certain materials possessed properties which he called acid, while others possessed other properties which he called bases. Between these two was a neutral area in which he called bases. Between these two was a neutral area in which the material showed neither acid or base properties and he termed this a neutral.

Now, just as in hot and cold, the words acid and base do not give us a scientific value that we can use. We needed a scale such that we all would be in exact agreement when we discussed the degree of acidity or basicity.

Some time ago a fellow named Sorensen developed such a scale and at the same time came up with the symbol pH.

Before we study the scale, let's take a moment to find out what makes one material an acid while another is a base.

1. An acid must have ionized (or free) hydrogen ions, H^+.

2. A base must have ionised (or free) hydroxyl ions OH^-.

3. The pH is directly related to the ratio of H+ to OH^-. If the H^+ is greater than OH^- the material is acid. If the OH^- is greater than the H^+, the material is a base. If equal amounts are present, the material is a neutral salt.

Back to the development of the scale. A molar solution of hydrochloric acid is about a 3.6% solution of the acid. Let's assign this solution a pH of 0.

A molar solution of sodium hydroxide (lye) is about a 4.0% solution of the base. Let's assign this solution a pH of 14.

If we now dilute the acid by adding one milliliter of acid to nine milliliters of pure water we will have a 1/10 molar solution. Let's assign this solution a pH of 1. In the same way, let's dilute the molar solution of sodium hydroxide by adding 1 milliliter to 9 milliliters of pure water. We'll assign this resultant solution a pH of 13.

Notice that with the acid, our 1/10th dilution increased the pH from 0 to 1, while the same dilution of the base reduced the pH from 14 to 13. Now here's the significant point - in both cases we were going toward 7. As we'll see shortly, a pH of 7 is the exact middle of the scale at which we have neither acid nor base, but neutral.

Continue to dilute both the acid and base by 1/10th and each time increase the number in the case of the acid and decrease the number in the case of the base.

The results look like this

Acid pH		Base pH	
1.0 molar	0	1.0 molar	14
0.1 molar	1	0.1 molar	13
0.01 molar	2	0.01 molar	12
0.001 molar	3	0.001 molar	11
0.0001 molar	4	0.0001 molar	10
0.00001 molar	5	0.00001 molar	9
0.000001 molar	6	0.000001 molar	8
0.0000001 molar	7	0.0000001 molar	7

Notice an interesting fact on the acid column - the number of decimal places exactly equals the pH!. Therefore a change of 1/10th or 10 times in acid concentration will make a change of 1 pH unit either up or down. The same holds true for bases. A change of either 1/10th or 10 times will change the pH by 1 either up or down.

The following is really not plain language and can be skipped without too great a loss.

> pH is really a number that is the negative logarithm of the molar concentration of H^+ (hydrogen ions).

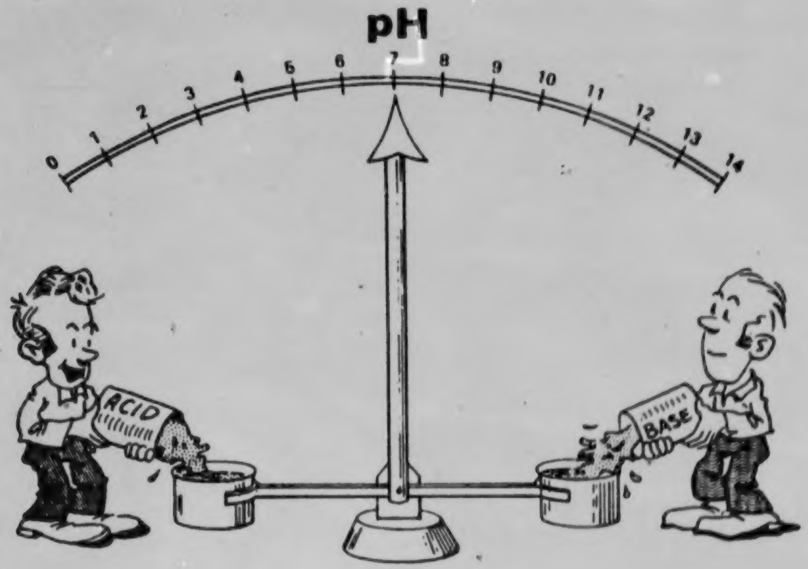
Another interesting fact from the chart: if we mix equal volumes of pH 3 (0.001 HCl) and pH 11 (0.001 NaOH), the resultant solution will be a pH of 7. The same applies to 0 and 14, 1 and 13, and so on. Equal amounts of H^+ and OH^- neutralize and result in a pH of 7.

A logical question at this point is "Do all acids and bases act the same?" Unfortunately they do not. The major difference comes in the amount of free H^+ or OH^- that is present in solution. For example, when you dissolve hydrochloric acid in water, all of the hydrogen ions H^+ are free in solution. The same is true of the sodium hydroxide and the hydroxyl ions OH^-. However if you dissolve acetic acid in water, all of the hydrogen ions H^+ are not free (only about 1.3% of the total H^+) and the pH is considerably higher than what you would expect for a molar solution of the acid. Thus, acetic acid (vinegar) is called a weak acid. A good example of a weak base is ammonium hydroxide (household ammonia).

The subject of ionization and dissociation will not be covered h[illegible]ust bear in mind that not all acids or bases are completely ionized.

S PAGE

Refer to the scale below to picture just how pH relates to acid and base. Adding acid on the left tips the scale pointer to the left and the number gets smaller (more acidic). Adding base on the right tips the scale pointer to the right and the number gets larger (more basic).



HOW DO WE MEASURE pH?

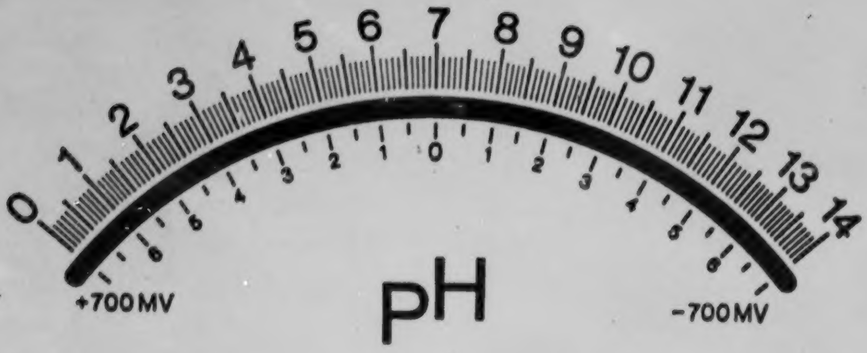
Although there are paper and dye indicators that change at different pH values, we shall be concerned only with pH meters since we are interested in a precise number that represents the true pH of our solution.

First, there is a pH probe (which we'll discuss shortly) which produces a voltage that can be directly related to the pH of the solution in which we place the probe. Secondly, there is an electronic circuit within the pH meter cabinet that receives the voltage from the probe and then presents it to the meter scale. This voltage developed at the probe will cause the meter pointer to move. The value of the number at which the pointer stops is the pH of the solution.

Let's say we had three separate solutions. No. 1 contains 0.001 molar hydrochloric acid. No. 2 contains pure water and No. 3 contains 0.0001 molar sodium hydroxide.

No. 1	No. 2	No. 3
____	____	____

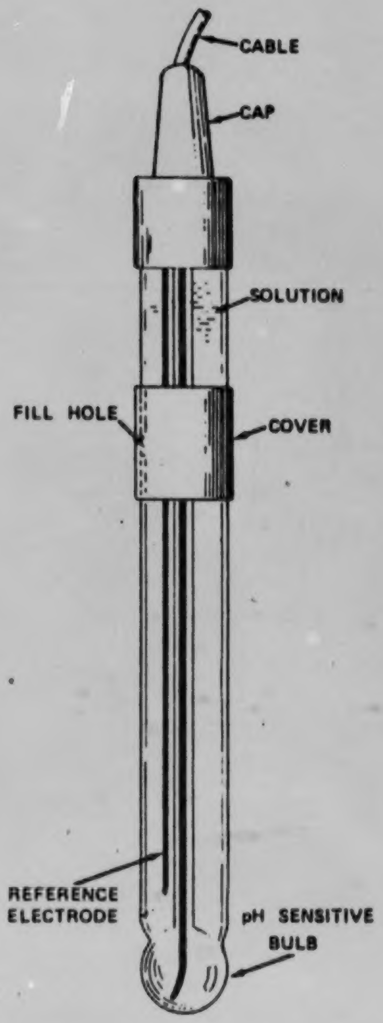
Fill in the blanks with the number that you would expect the pH meter to read as you move the pointer from solution to solution. (If you wrote in No. 1 pH 3, No. 2 pH 7 and No. 3 pH 10, you get 100.)



There are ten small divisions between 7 and 6. Each small division is equal to 0.1 pH units. Thus if the meter pointer came to rest at the fourth small division to the right of the number 6, the pH would be 6.4. If the pointer came to rest between the small divisions, one can estimate and write the value as 6.45.

The shiny strip just below the scale is called mirror-backing. Its purpose is to allow you to look at the pointer and if you see a reflection of the pointer in the mirror you move your head slightly until you cannot see the reflection. This means that you are looking directly at the pointer and therefore the value you read is correct. You can readily see that if two people read the same meter but from different angles, they would report different values. The mirror backed scale helps us all to look at the same angle.

The probe can be thought of as a battery whose voltage changes as the pH of the solution in which it is inserted changes. It consists of two parts (in fact many pH measurements are made with two separate probes), first the hydrogen sensitive glass bulb and second the reference electrode. The special glass of the bulb has the ability to pass H^+. This ability then allows the H^+ inside the bulb to compare themselves with those outside the bulb and to develop a voltage that is related to this difference. This bulb then is a half-cell and will need a companion reference to function.



On the drawing notice that just above the bulb is a reference electrode frit. This is actually a small opening in the glass through which the inside filler solution can very slowly leak out. Now the relationship between the reference electrode and the solution also produces a voltage and this is the other half cell. Together the pH sensitive bulb and the reference electrode constitute the complete probe.

The value of voltage produced by the probe is fortunately a linear function of the pH. For example, at pH 7.00, the probe produces zero volts while at pH 6.00 it produces +0.06 or +60 millivolts. Notice the plus polarity mark; if the voltage were of minus polarity, the meter pointer would go to the right to a pH reading of 8.00.

Generally, a probe will produce about 60 millivolts for each change of 1 pH unit. Thus, a probe voltage of +300 millivolts would cause the meter to read pH 2.00 (+300 ÷ 60 = 5 units, 7 - 2 = 2).

Since the pH meter and probe are both electronic-type devices, you might like to have some standard pH solution that you can measure and be sure that everything is calibrated correctly. These solutions are available and they are called standard buffers. You'll soon find that a buffer is a vital part of all pH measurements.

A buffer is a solution of a particular pH that has the ability to resist change in pH. (By the way, buffers in our blood system are what keep us alive and healthy.)

Here's an experiment to demonstrate the ability of buffers to resist change. Take two glasses and in one place distilled water. In the second, place a buffer near pH 7.00. Next measure and record the pH at the start. They should both be near 7.00. Now, add one drop of concentrated hydrochloric acid to each glass. Stir, and again record the pH. Do this for several drops with a measurement between each drop. You should find the distilled water drop in pH with each drop. The buffer, on the other hand, should resist change until you exceed its buffering capability.

We won't go into why buffers act this way. Suffice to say it has to do with a phenomenom called common-ion effect. Look it up in your chemistry text.

Temperature is also a very important factor in pH measurements and here's why. We said earlier that the voltage output of a probe was about 60 millivolts per pH unit. However we didn't confuse you by telling you that this is true only at 25°C (room temperature). If you reduce the temperature to 0°C, the probe will provide only about 54 millivolts. And if you increase the temperature to 100°C the probe will provide about 70 millivolts per pH unit.

You can easily see that you must have some way of correcting for this, or make all of your measurements at 25°C.

Let's re-state this situation as a problem and see if we can get a clear understanding of what's happening.

Problem: How to make a pH meter read the correct value of a buffer or sample solution at all temperatures.

Probe output for each pH unit

0°C	54 millivolts
25°C	60 millivolts
50°C	64 millivolts
75°C	68 millivolts
100°C	70 millivolts

In looking at this table we realize that our pH meter must be able to move one pH unit with 54 millivolts input or 70 millivolts or any value in between. Most meters can do this.

Solution: Let us put a sensitive attenuator in the meter and adjust it first to work with 54 millivolts to move 1 pH unit. We'll have a knob on this attenuator and mark the point at which 54 millivolts equalled 1 pH unit. Next, let's move the attenuator to a position where it requires 70 millivolts to move the pointer 1 pH unit. (By the way, we can apply these voltages from a voltage source, and do not require a probe.) In the same way we can set 60 millivolts and so forth. When we are finished we'll have a knob control in which we can set the proper sensitivity of the meter to correspond to the temperature of the sample solution.

Finally, since we are really concerned only with pH (and not millivolts), let's erase the 54 millivolts and write instead 0°C. For the 60 millivolts write 25°C and so forth around the circle.

At this point let's review a little and then we'll make a bona fide pH measurement.

First we discussed the meter and its mirror-backed scale. Next we found that probes are really two-piece affairs (reference and pH sensitive bulb) that produce millivolt outputs and vary with pH. Then the important buffers were described and finally we found that temperature played a vital role in accurate measurements.

To pass the course, we must successfully use this information and produce an accurate pH measurement. Assume for the moment that someone has just handed us a bottle containing a watery solution and asked "What is the pH of this?" Here's how to proceed.

1. Be sure your pH meter has been calibrated and is operating correctly as described in your instruction manual.

2. Turn on the pH meter and allow a few minutes warmup.

3. Set the temperature dial to the temperature of the sample (we will assume that the sample and buffers are at room temperature).

4. Place the pH probe in a 7.00 buffer and position the pointer with the Set knob for a reading of 7.00.

5. Remove the probe, rinse, wipe, and then immerse it in the unknown.

6. Note the reading. If it is near 7.00, record the exact value and the measurement is complete.

7. If it is below 6.00 or above 8.00, then select a buffer that is closest to the unknown. Place the probe in this buffer after a rinse and wipe. Use the Set control to make the meter read exactly the value of the buffer.

8. Remove the probe, rinse, wipe and place in the unknown. Note the reading. The accuracy of this reading will depend primarily on the accuracy of the standard buffer. The closer the buffer value to the unknown, the more accurate the reading.

NOTE: If the unknown in this example had read a value above pH 12.00, you should substitute a special probe that is designed for pH measurement in highly alkaline solution. Most general purpose probes begin to lose accuracy above pH 12.00 due to sodium ion.

Before we end this section on pH measurements, let's discuss another problem and find a solution.

Problem: You are required to measure a series of solutions that may differ by only 0.1 pH units.

Discussion: The meter we described had 10 small marks between each pH unit and each of these marks was 0.1 pH units. In the problem above, it would be very difficult to see the small differences on that meter scale.

Solution: We expand the scale ten times. Thus we use an expanded scale meter and each pH number is now equal to 0.1 pH unit and each small mark equals 0.01 pH unit. Now all we need is a buffer in the vicinity of our unknown and we ca easily and accurately make the required measurements.

Of course, we could also have used a digital readout pH meter and read the results directly.

pH TEST*

The pH scale measures the intensity of acidity and alkalinity. The pH scale runs from 0 to 14.

---Water with a 7.0 pH level is neutral. It is neither acidic or alkaline.
---Water containing acid will have a pH level below 7.0. The stronger the intensity of the acid, the lower the pH will be. A pH of 0 means the water is very acidic.
---Alkalies raise the pH level above 7.0. A pH of 14 means the water is intensely alkaline.

The pH level is important to chlorination, coagulation, softening, and corrosion control.

(a) The pH of water affects the disinfecting action of chlorine, particularily combined available chlorine residual. For example: at pH 6.5 and a temperature of 70°F, 0.3 ppm of combined residual causes 100% bacterial kill. With the same temperature at pH 7.0 the combined residual must be increased to 0.6 ppm to accomplish the same degree of bacterial kill.

(b) Several chemicals used to aid coagulation are either acid or alkaline compounds and are used to adjust the pH for best coagulation.

(c) In the softening process, magnesium and calcium must be precipitated. Different pH levels are used for each: about pH 9.5 for calcium carbonate and pH 10.5 for magnesium hydroxide.

(d) The pH level is important to the control of corrosion. The pH level controls the solubility and precipitation of substances such as ferric hydroxide and calcium carbonate. The pH governs the single electrode potential of a metal exposed to it - influencing the types of reactions that occur at the cathode and the anode.

*Source: Adapted from Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, Chapter XIX, pp. 8-9.

pH TEST

There are two methods for determining the pH value: colorimetrically and electrometrically. The electric pH meter is far more accurate than the colorimetric method.

Colorimetric pH Determination

Equipment and Reagents:

pH comparator
Standard color discs for comparator
Dropper pipettes
Glass tubes or cells to fit comparator
Indicator solutions

The indicator solution and the pH range of each disc is listed below. Two or more of these are usually needed.

Bromcresol green . . . . . pH range 4.0 to 5.6
Chlorophenol red . . . . . pH range 5.2 to 6.8
Bromthymol blue . . . . . pH range 6.0 to 7.6
Cresol red . . . . pH range 7.2 to 8.8
Thymol blue . . . . pH range 8.0 to 9.6

(The indicator solution should be obtained from the manufacturer of the disc it is to be used with. Give the manufacturer the disc number when ordering.)

There are several different types of comparator sets available. The procedure for making the pH test will vary slightly with the different comparators.

Sampling:

Samples for pH determination must be collected carefully. Small quantities of foreign material can affect the results. Be sure your hands are clean. If you have handled lime or alum there may be enough chemical on your hands to affect the results if they come in contact with the sample.

Procedure:

(1) Add the required amount of indicator *(exactly as specified by the manufacturer)* to one of the two comparator tubes.

(2) Fill each comparator tube to the mark with the sample. *(If the indicator is added first, the filling of the tube will cause thorough mixing. This eliminates the need for placing a thumb or finger over the tube in order to shake it. Lime, alum, perspiration, or other contaminants may change the pH value.)*

(3) Place the tube containing the indicator in the inner slot of the comparator. Place the tube with no indicator in the outer slot. *(The tube with no indicator, sample only, is to compensate for any color which may be in the water.)*

(4) Look through the eyepiece while holding the comparator up to daylight (not sunlight) and rotate the disc until the colors seen through the eyepiece match. Record the pH value from the exposed number.

(5) Results are reported as pH value and are read directly from the comparator or pH meter. The accuracy of colorimetric pH determinations can be vastly improved by dechlorinating the sample with sodium thiosulfate crystals (or solution) before the pH indicator is applied. *Free chlorine residuals of almost any magnitude bleach the pH indicator colors so quickly that (unless the chlorine is destroyed) the readings can be as much as one full pH value in error.*

Electronic pH Determination

Equipment and Reagents:

Electric pH meter complete with electrodes
2 Beakers, 100 - 200 ml capacity
Buffer solution, pH 6.86 (A standard)
(Note: An additional solution, pH 4.01 would be useful.)
Distilled water

Procedure:

(1) Measure the temperature of the sample. Set this temperature on the temperature dial of the meter according to the manufacturer's instruction booklet. Turn the unit on according to the manufacturer's instruction booklet. Allow the prescribed "warm up" time. *(Make sure the electrodes are immersed in water when the machine is not in use.)*

(2) Make sure that the meter selector switch is on "Standby". Raise the electrodes and rinse with distilled water. Wipe electrodes dry with very soft tissue. Repeat this procedure every time the unit is used.

(3) Pour a sufficient amount of the buffer solution into a clean beaker and calibrate the pH meter with each buffer solution according to the manufacturer's instructions. Calibration with one buffer solution is satisfactory if the second solution is not available.

(4) Repeat Step (2).

(5) Determine the pH of the sample according to the manufacturer's instructions. *(Note: When the meter is left on but is not being used, most manufacturers require that the meter be left on "Standby".)*

(Chemistry/Bacteriology # 16)
Evaluation Form
EVALUATOR'S COPY
Part I

Operator ______________________

Evaluator ______________________

Date __________ — ______________

Directions: Write the word "True" or "False" next to each item in the space provided. If you answer "False" write the correct answer on the lines under the question.

1. *False* The higher the pH value, the higher the acidity of the water.
pH value above 7 indicates high alkalinity; greater acidity lowers the pH value.

2. *True* A high pH level (7 or more) helps to eliminate corrosion problems in the distribution system.

3. *True* The pH level can affect coagulation.

4. *False* The pH level has no affect on chlorination.
The pH level combined with the temperature affects the disinfecting action of chlorine.

5. *False* The pH test measures the total amount of acid and alkali present in water.
The pH test yields a reading of the relative amounts of acid and alkali. The strength, not the amount of acid or alkali, has the greatest effect on the pH level.

(Chemistry/Bacteriology #17)
Competency Sheet

COMPETENCY: Describe the process of coagulation and the purpose it serves in water treatment.

Objective: Describe the process of coagulation.

Objective: Explain where in the treatment process coagulation takes place.

Objective: Describe "floc", how it is formed, and the purpose it serves.

RESOURCES:

(a) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XIII. *-Gold Appendix-*

(b) Manual of Instruction for Water Treatment Plant Operators, New York State Department of Health, Health Education Service, P.O. Box 7283, Albany, New York 12224, Chapter 7. *-Pink Appendix-*

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about the purpose and process of coagulation.

(b) Observe the coagulation process as it takes place at your treatment plant. Make note of the type and amount of chemicals used. How are the chemicals measured? How are they added to the water?

(c) Describe "floc". What causes floc to form? What purpose does floc serve?

EVALUATION ACTIVITY: Complete Evaluation Form #17. It contains questions about the coagulation process and the purpose it serves. You may not use references.

(Chemistry/Bacteriology #17)
Evaluation Form
EVALUATOR'S COPY
Page 1 of 2

Operator ____________________

Evaluator ____________________

Date ____________________

Directions: Write the word "True" or "False" in the space provided next to each item. If you answer "False" write the correct answer on the line under the question.

1. *True* Aluminum sulfate and ferric sulfate can be used as coagulants.

2. *False* During sedimentation floc is formed.
Floc is formed during coagulation. Floc should settle out during sedimentation.

3. *True* A "coagulant" is a chemical used to help form a good floc.

4. *True* Usually, the more turbidity there is, the more alum is necessary to form a good floc.

5. *True* After-floc may form in the clear well.
Note: Coagulant in suspension may pass through filters and settle out in reservoirs or mains. After-floc may form in clear water if water passes filters before post-lime is added for corrective treatment.

6. *True* There is at least one pH zone for every water which produces the best floc with the least chemicals.

7. *False* Chemicals used for coagulation should not be rapidly mixed with the water.
Rapid mixing distributes the chemical(s) throughout the water and helps form a good floc.

8. *False* Very fine or fluffy floc settles better than coarse, heavy floc.
Fine floc does not settle well. Fluffy floc hangs near the surface and may travel the entire length of the settling basin.

9. *True* Coagulation causes electrical and chemical reactions that cause the fine materials in water to cluster together.

10. *True* A "jar test" helps to determine the most economical amount of chemicals to use for proper coagulation.

11. *True* Any chemical (such as lime, soda ash, chlorine, fluoride) should be added before alum is added.

12. *False* Coagulation usually takes place after filtration.
Coagulation precedes filtration. It helps take some of the stress off the filter.

(Chemistry/Bacteriology #18)
Competency Sheet
Class II*

COMPETENCY: Perform a "jar test" to determine the chemical dosage and pH level at which water in your plant flocs best.

Objective: Describe the purpose and procedures for performing a "jar test".

Objective: Demonstrate the ability to perform a "jar test" accurately.

RESOURCES:

(a) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XIII. *-Gold Appendix-* See section titled "Jar or Coagulation Test".

(b) Laboratory Equipment and Supply Requirement For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about jar tests and how they are performed.

(b) Practice the procedures for the jar test. If possible, have an experienced operator observe you while you perform the test to help you if you have any questions. Remember to operate the stirring equipment similar to the mixing and/or flocculating conditions in your plant. It is important to *record in writing* the dose of each stock solution added to each beaker. *The smallest amount of chemicals which produces the best settling floc in your jar test will indicate the rate at which the incoming raw water should be treated.*

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #18. It contains questions about the purpose and procedures for performing a "jar test". You may not use references.

(b) Perform a "jar test" using the proper equipment and procedures. Indicate the most economical chemical dosage for coagulation/flocculation. You may use references during this part of the evaluation.

**This competency is required for Class II operators and persons applying for Class II certification. Not required for Class I.*

(Chemistry/Bacteriology #18)
Evaluation Form
EVALUATOR'S COPY

Operator ____________________

Evaluator ____________________

Date ____________________

Directions: Write the word "True" or "False" in the space provided next to each item. If you answer "False" write the correct answer on the lines below the question.

1. *False* The purpose of a "jar test" is to determine how many chemicals are present in the raw water.
The purpose is to identify the most economical chemical treatment which will provide effective coagulation/flocculation.

2. *False* The pH level does not affect the coagulation process.
pH influences coagulation. Note: All coagulants do not react similarly at the same pH.

3. *True* The temperature of the water affects coagulation.

4. *True* The speed of stirring during "jar tests" should be similar to the speed of the flocculators used in the plant.

5. *False* Color is usually removed when water is highly alkaline.
Color is more often readily removed at an acid pH.

6. *True* Turbidity is usually coagulated under alkaline conditions.

(Chemistry/Bacteriology #19)
Competency Sheet

COMPETENCY: Determine the pH level most appropriate for controlling corrosion in your treatment area.

Objective: Describe problems caused by corrosion.

Objective: Describe how the pH level affects corrosion.

Objective: Determine the pH level that best controls corrosion in your treatment area.

RESOURCES:

(a) Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual No. 85-13, June 15, 1968, pp. 209-228, 235-36. *-Buff Appendix-*

(b) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XX. *-Green Appendix-*

(c) "Symptoms of Corrosion" Resource Sheet.

(d) Resource Chart (Baylis Curve).

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about corrosion and how it is affected by the pH level.

(b) Describe various ways to diagnose corrosion. Check to see whether hydraulic surveys are made for your plant to determine the carrying capacity of the water mains. Reductions in carrying capacity caused by corrosion or scale formation can be detected using these surveys.

What checks for corrosion are done for your distribution system?

(c) Describe how the pH level of finished water affects corrosion in the distribution system.

(d) Using the chart provided, determine the pH level most appropriate for controlling corrosion in your treatment area distribution system.

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #19. It contains questions about problems caused by corrosion and the effects of the pH level on corrosion. You may not use references.

(b) Determine the pH level most appropriate for controlling corrosion in your treatment area distribution system. You may use references.

SYMPTOMS OF CORROSION*

1. "Red" water.

2. "Pitting" of iron results from the solution of iron.

3. Blue/green stains on enamel fixtures indicates corrosion of copper. Soft waters high in carbon dioxide are generally corrosive toward copper.

4. Corrosion of aluminum produces a white aluminum oxide deposit on metal.

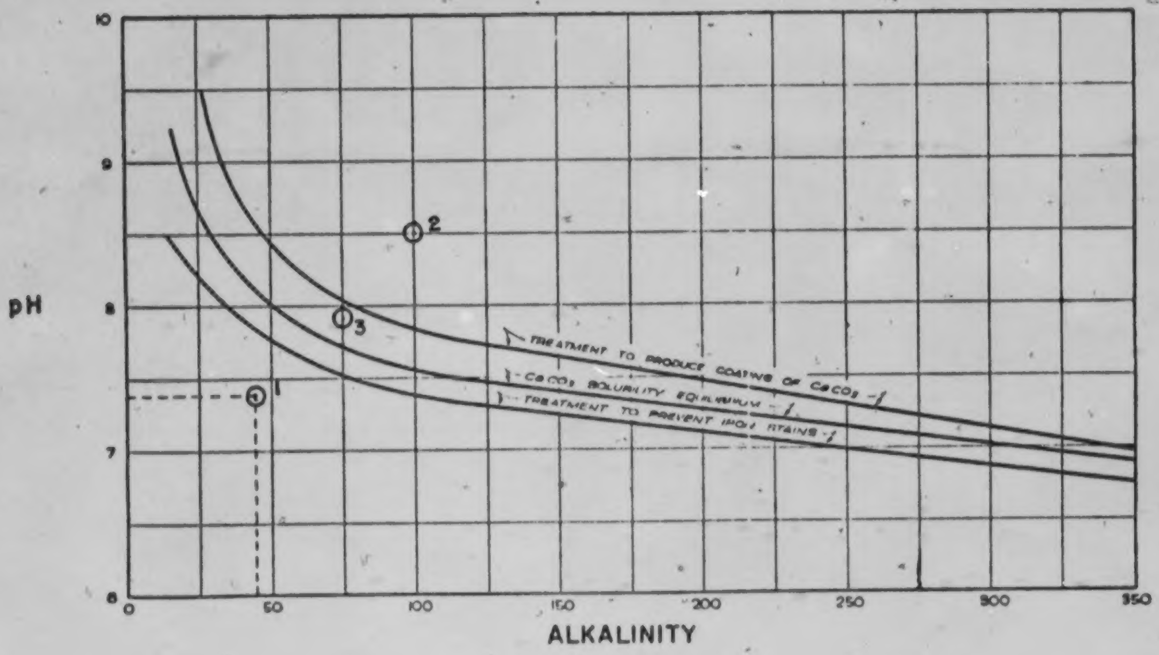
5. Soft waters high in carbon dioxide may corrode lead pipe forming toxic (poisonous) lead salts.

6. Corrosion may effect the operation of heating and air conditioning equipment.

*Source: Adapted from Manual of Instruction for Water Treatment Plant Operators, New York State Department of Health, Health Education Service, P.O. Box 7283, Albany, New York 12224, p. 197.

(Chemistry/Bacteriology #19)
Resource Chart

"BAYLIS CURVE"



RELATION OF pH AND ALKALINITY TO THE PROPER TREATMENT OF WATER TO PREVENT CORROSION
The calcium carbonate solubility curve shown is based upon calcium carbonate in distilled water at 22°C It is approximately correct for natural waters low in soluble salts but should be checked when used on waters containing considerable dissolved compounds.

(Chemistry/Bacteriology #19)
Evaluation Form
EVALUATOR'S COPY
Part I

Operator ______________________

Evaluator ______________________

Date ______________________

Directions: Write the word "True" or "False" in the space provided next to each item. If you answer "False" write the correct answer on the lines below the question.

1. *True* Corrosion often causes water to have a reddish color.

2. *True* Blue/green stains on enamel bathroom fixtures may indicate corrosion of copper pipe.

3. *True* The corrosion of lead pipe causes toxic (poisonous) lead salts to form.

4. *False* Corrosion is usually more rapid when the pH level is high (alkaline).
Corrosion is generally more rapid in low pH (acid) water than in alkaline solutions.

5. *True* Rust always forms when the pH level is between 4.5 and 9.0.
Note: This is the pH range of most natural waters.

(Chemistry/Bacteriology # 20)
Competency Sheet

COMPETENCY: Perform an alkalinity test on your raw and finished water.

Objective: Describe the purpose for performing alkalinity tests.

Objective: Properly perform alkalinity tests on your raw and finished water.

RESOURCES:

(a) "Alkalinity" Resource Sheet.

(b) Laboratory Equipment and Supply Requirements For a Small Water Treatment Plant Providing Complete Treatment *-Salmon Appendix-*

LEARNING ACTIVITIES:

(a) Read the Resource Sheet and Appendix to find information about alkalinity tests.

(b) Practice performing tests for alkalinity. If possible, have an experienced operator observe you to help answer any questions you may have. Once you have determined alkalinity, describe how coagulation will be affected.

EVALUATION ACTIVITIES:

(a) Complete Evaluation Form #20. It contains questions about alkalinity. You may not use references.

(b) Perform an alkalinity test while your evaluator observes you. Determine the total alkalinity in terms of mg/l as calcium carbonate. You may use references while you perform the test.

"ALKALINITY"*

Purpose of Test

Many of the chemicals used in treating water can cause a change in its alkalinity. The most pronounced changes are caused by coagulants and by the softening chemicals lime and soda ash. This alkalinity test is intended to provide results for use in calculating the chemical dosages needed for coagulation and softening. Total alkalinity must also be determined in the course of the calcium carbonate stability test and in estimating carbonate hardness.

Warning

This method is suitable for the titration of waters containing hydroxide, carbonate, or bicarbonate alkalinity. The water should be free of color and turbidity (which might obscure and affect the indicator response). The residual chlorine content of the water should not be greater than 1.8 mg/l.

Equipment

---A 25 ml buret and support.
---A 100 ml graduated cylinder (or appropriate volumetric pipets for measuring the sample).
---Two or more 250 ml flasks or porcelain casseroles.
---Two or more stirring rods.
---Three dropping pipets of 0.5-1 ml capacity for dispensing sodium thiosulfate, phenolphthalein indicator, and methyl orange indicator solutions.

Reagents

---Phenolphthalein indicator solution. Weigh 0.5 g phenolphthalein disodium salt powder and dissolve in 100 ml distilled water.
---Methyl Orange indicator solution. Weigh 0.05 g methyl orange indicator powder and dissolve in 100 ml distilled water. The powder may take quite a bit of time and stirring to dissolve at room temperature.
---Sulfuric acid titrant, 0.0200N H_2SO_4. This solution requires some skill to prepare, standardize, and adjust to exactly 0.0200N. The 0.0200N solution can be purchased from reputable supply houses.
---Sodium thiosulfate solution, 0.1N *(not required if the water contains no residual chlorine)*. Weigh 2.5 g sodium thiosulfate, $Na_2S_2O_3 \cdot 5H_2O$, and dissolve in 100 ml distilled water.

*Source: Adapted from Simplified Procedures for Water Examination, American Water Works Association, AWWA No. M12, 1975, pp. 16-19.

Procedure

A. Fill the buret with sulfuric acid titrant. Record the liquid level in the buret by reading at the bottom of the meniscus. Be sure to guard against a leaky stopcock which results in the loss of titrant if left standing.

B. Measure the correct sample volume for the indicated alkalinity ranges:

"TABLE I"

Sample Volume ml	Alkalinity Range mg/l as $CaCO_3$
100	0 - 250
50	251 - 500
25	501 - 1,000

Example: If the alkalinity falls within range of 0-250 mg/l as $CaCO_3$ (calcium carbonate), take a 100 ml sample.

Place equal volumes of the sample into two 250 ml flasks (or porcelain casseroles). One will be used as a blank for color comparison.

C. *If necessary, remove the residual chlorine by adding 1 drop (0.05 ml) of sodium thiosulfate solution to each flask (or casserole). Mix.*

D. Add two drops of phenolphthalein indicator solution to one flask and mix. If the sample turns pink, carbonate or hydroxide is present; go on to step "E". If the sample remains colorless, the water contains bicarbonate or is acid; skip steps "E" and "F". Go on to step "G".

E. If the sample turns pink, gradually add sulfuric acid titrant from the buret. Shake the flask constantly until the pink just disappears. Use the flask without any phenolphthalein indicator as the color comparison blank.

F. Read the new buret level at the bottom of the meniscus. Calculate the volume of acid used by subtracting the initial buret reading (Step A) from the present reading.

G. Calculate the phenolphthalein alkalinity in terms of mg/l as calcium carbonate: multiply the result from Step "F" by the correct factor.

For a sample volume of 100 ml, multiply milliliters of H_2SO_4 by 10; for a 50 ml sample, multiply by 20; for a 25 ml sample, multiply by 40.

"TABLE II"

Calculation of Alkalinity from Titration Results[*]

	Alkalinity - mg/l as $CaCO_3$		
Titration Result	Hydroxide	Carbonate	Bicarbonate
P = 0	0	0	T
P less than ½T	0	2P	T - 2P
P = ½T	0	2P	0
P more than ½T	2P - T	2T - 2P	0
P = T		0	0

[*] *P = phenolphthalein alkalinity (mg/l); T = total alkalinity (mg/l)*

H. Add 2 drops (0.1 ml) of methyl orange indicator solution to both of the flasks containing the water sample.

I. Titrate again with small volumes of sulfuric acid until the yellow color just begins to change to an orange. Continually compare the color of the sample with that of the blank against a white background. At the endpoint a slight difference should be noticeable between the color of the comparison blank and the titrated sample. The comparison blank remains constant yellow-the titrated sample gradually turns a faint orange. The color change takes place within the span of 2 to 4 drops of sulfuric acid.

J. Read the buret again. Calculate the total volume of acid used in both the phenolphthalein titration and the methyl orange titration. Multiply this total by the correct factor given in Step "G". *The result is the total alkalinity* (also called methyl orange alkalinity) *in terms of mg/l as calcium carbonate.*

Practical Applications of Alkalinity Titrations

A. The phenolphthalein alkalinity (P) and the total alkalinity (T) can be used to calculate the hydroxide, carbonate, and bicarbonate in terms of calcium carbonate. The titration to methyl orange (T) measures all of the alkalinity due to bicarbonate, carbonate, and hydroxide. The titration to phenolphthalein (P) measures all of the hydroxide and exactly ½ of the carbonate, but none of the bicarbonate. Therefore,

   1) Bicarbonate is present when P is less than ½ of T.
   2) Carbonate is present when P is not zero but is less than T.

Table II shows how to calculate the various types of alkalinity. The calculations are based on the following assumptions:

---Bicarbonate and carbonate alkalinities may exist in the same sample.
---Carbonate and hydroxide cannot coexist.
---Carbon dioxide may be present when P is absent.
---Conversely: The existence of P means the absence of carbon dioxide in the water.
---A pH of 8.3 represents the borderline of carbonate alkalinity.
---Since mineral acids are not normal constituents of water, their presence indicates pollution.

3) Hydroxide is present when P is more than ½ of T.

B. Hydroxide or caustic alkalinity in water undergoing softening treatment indicates the presence of excess lime. This may be in the range of 10-50 mg/l. The alkalinity can be reduced to a suitable level by carbonation.

C. Soda ash (sodium carbonate) is added to some soft waters to provide sufficient carbonate alkalinity to improve alum coagulation. The alkalinity titrations afford a means of properly controlling the soda ash dosage.

(Chemistry/Bacteriology #20)
Evaluation Form
EVALUATOR'S COPY
Part I
Page 1 of 2

Operator ______________________________

Evaluator ______________________________

Date ______________________________

Directions: Write the word "True" or "False" in the space provided next to each item. If you answer "False" write the correct answer on the lines under the question.

1. *False* Residual chlorine will not affect the results of the alkalinity test.

Residual chlorine should not exceed 1.8 mg/l. Residual chlorine can be removed with sodium thiosulfate solution.

2. *True* Alkalinity tests provide information needed for calculating chemical dosages for coagulation.

3. *True* Color and/or turbidity may affect the results of the alkalinity test.

4. *True* The presence of mineral acids in water indicates pollution.

5. *False* Bicarbonate and carbonate alkalinities cannot exist together in the same sample.

It is not uncommon for bicarbonate and carbonate alkalinities to exist in the same sample.

6. *False* A pH of about 4.0 represents the borderline of carbonate alkalinity.

A pH of 8.3 represents the borderline of carbonate alkalinity. Answers of pH 8.0-8.5 are acceptable.

(Chemistry/Bacteriology #21)
Competency Sheet
Class II*

COMPETENCY: Determine the rate of feed for each chemical dry feeder in your treatment plant and accurately measure the correct amount of dry chemical and add it to each dry feeder in your plant.**

Objective: Identify the dry feeders in your plant and the chemical being fed by each one.

Objective: Determine the appropriate rate of feed for each chemical dry feeder in your treatment plant.

Objective: Accurately measure the correct amount of dry chemical and add it to each dry feeder in your plant.

RESOURCES:

(a) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XIII. *-Gold Appendix-* See section titled "Dry Chemical Feeders".

(b) "Chemical Dry Feeders" Resource Sheet.

(c) Manufacturer's operators' manual for each chemical dry feeder in your plant.

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about chemical dry feeders.

(b) Take a "tour" of your treatment plant. Identify each dry feeder and the chemical it is feeding. Obtain and review the manufacturer's operators' manual for each dry feeder.

(c) Under supervision of an experienced operator practice measuring and loading the chemicals for each dry feeder in your treatment plant. Be sure to follow the manufacturer's instructions for each dry feeder.

EVALUATION ACTIVITY: While your evaluator observes, accurately measure the correct amount of dry chemical and add it to each dry feeder in your treatment plant. Adjust each feeder to deliver the chemical at an appropriate rate.

**This competency is required for Class II operators and persons applying for Class II certification. Not required for Class I.*

**If there are no chemical dry feeders in your treatment plant, make arrangements with your evaluator to visit a plant with dry feeders.

"CHEMICAL DRY FEEDERS"*

Volumetric Dry Feeders

a. *Clean and Inspect.* Clean feeder, feeder mechanism, and feeder surroundings. Remove spilled chemical and chemical-dust accumulations with vacuum cleaner or brush. Look for oil drips, wiring defects, and deterioration. Observe general performance of feeder. Note and investigate unusual noises. See that orifice, knife edges, scrapers, shakers, and opening are free of chemical accumulations. Probe solution tank for sediment or undissolved chemical and make necessary corrections. Be sure that hopper is filled without chemical spillage on working parts of feeder. If chemical feeder is out of service, see that feeder is empty of chemical and that condensation is not causing corrosion and deterioration.

b. *Inspect for Loose Bolts and Cracks.* Carefully wipe all parts of feeder. Inspect for loose bolts, cracks, defective parts, and leaks.

c. *Test Calibration.* Make several calibration tests, using range of feed in which chemical feeder operates most frequently, to check accuracy of rate-setting devices, charts, or tables.

(1) *Test with Feeder Not On Weighing Scale.* With chemical feeder delivering uniformly after a change in rate of feed, catch the chemical discharged in a pan for a specific period of time. Weigh material caught and calculate feed in pounds per hour. Compare with indicated rate setting or chart. Be careful and accurate in timing, weighing, and collecting material. Perform each test several times to check results.

(2) *Test with Feeder Having Weighing Scale.* With feeder weighing scale operating, check rate of feed by stopping feeder, balancing scales, and then running feeder for a specific period of time. Stop feeder and rebalance scale. Subtract this reading from previous scale reading to determine amount of chemicals that passed through hopper in the given time. Calculate feed in pounds per hours and compare to rate setting. Do this for several rate settings in the operating range. Make any necessary adjustments.

d. *Service Drive Mechanism and Moving Parts.* Follow manufacturer's instructions in servicing and lubricating mechanism and moving parts.

e. *Overhaul.* Overhaul chemical feeder, clean, and paint interior and exterior. Clean and repair solution and makeup tanks. Service drive-mechanism bearings and other mechanical parts.

*Source: Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual #85-13, June 15, 1968, pp. 418-422.

Gravimetric Belt Type Dry Feeders

a. *Clean, Check, and Test Balance.* Clean feeder, feeder mechanism, and surroundings. Empty feed belt, clean off accumulations, and test scale balance with zero load. Make necessary adjustments to correct tare-weight balance. Check for mechanical and electrical defects and for proper operation. Look for oil drips and wiring defects. Observe general performance and investigate unusual noises. Make sure chemical is not accumulating on inner roll of belt or adhering to belt or sidewalls in a manner which may cause too high a tension or tear.

(1) *Check Solution Tank.* Probe solution tank for sediment or undissolved chemical. Clean and improve dissolving conditions if necessary.

(2) *Check Feeder out of Service.* If chemical feeder is out of service, see that feeder belt and hoppers are empty of chemical and that condensation is not causing deterioration.

b. *Inspect for Loose Bolts and Defective Parts.* Carefully wipe clean all parts of feeder, inspecting for and repairing loose bolts, leaks, defective parts, and so on. Check belt tension.

c. *Overhaul Feeder.* Remove feeder-belt assembly, motors, belts, and the like wherever practical and overhaul feeder completely. Clean and paint entire structure. Clean scales, and edges, and test for sensitivity. Lubricate all mechanical parts, and make any necessary repairs.

d. *Make Delivery Test.* Test delivery of feeder to see that its actual output is equal to weight indication on the totalizer or integration device.

(1) Clean belt and feeder and then balance scale. Return feeder to service and operate until chemical forms an even, uniform layer on feeder belt and scale remains in close balance. See that poise on scale beam indicates proper load on belt.

(2) *Stop Motor.* See that scale with its load of material moves freely and is in exact balance. If it is not, add or remove material from rear end of load on belt until exact balance is obtained.

(3) Disengage or stop operation of mechanism which feeds chemical to belt. Read belt revolution counter or weight integrator. Then operate feeder belt until counter or weight integrator indicates that exactly a predetermined number of pounds of chemical, about two-thirds of belt loading, have been fed. Stop feeder instantly by turning off motor switch.

(4) Again see that scale can move freely and then move poise to bring scale back to exact balance. Note scale reading. The difference between this scale reading and initial scale reading gives exact weight of material which has been delivered.

(5) If this reading does not coincide with number of pounds as shown by revolution counter, adjust as follows:

(a) If actual delivery as shown by front scale beam is greater than indicated by the revolution counter, move poise on scale beam to a lower figure so belt loading is less and repeat test.

(b) If actual delivery is less than indicated by revolution counter, move scale poise to higher reading on scale beam so belt loading is greater and repeat test.

(6) Continue testing until a belt loading is found which results in delivery of material corresponding exactly with reading of revolution counter. When this location is determined, set and clamp poise on scale beam at this position, and record setting for future reference. Thereafter when feeding same sort of material, use same belt loading. Recalibrate when using a different sort of material which has a different degree of slope at which material falls off front end of feeder belt.

Note: For certain other belt type feeders, a delivery test can only be made by catching the chemical delivered during an exact test period and weighing the catch. Actual delivery rate of feeder can then be calculated from these data.

Loss-in-weight Type Gravimetric Feeder

a. *Clean and Check.* Clean feeder, feeder mechanism, and surroundings. See that scale is sensitive to small changes in weight. Whenever feeder is completely empty, determine tare weight or see if scale indicates zero weight. Look for oil dips and wiring defects. Observe general performance of feeder. Note and investigate unusual noises. See that chemical is not accumulating on parts of the feeder, changing tare weight of scale. Probe solution tank for sediment or undissolved chemical. Clean tank and improve dissolving conditions if necessary. If chemical feeder is out of service, see that hopper and feed mechanism is empty and that condensation is not causing corrosion or deterioration.

b. *Inspect for Loose Bolts and Defective Parts.* Carefully wipe clean all parts of feeder, inspecting for and repairing loose bolts, leaks, defective parts, and the like.

c. *Service Drive Mechanism and Moving Parts.* Service drive mechanism and moving parts according to manufacturer's instructions.

d. *Overhaul Feeder and Scale.* Dissassemble feeder and clean and inspect all parts for wear or deterioration. Clean and paint feeder inside and out where necessary. Make necessary repairs and properly lubricate all mechanical parts. Service motors and drive mechanism.

Continuous Lime Slakers

a. *Check Dust and Vapor-removal Equipment.* For clean, trouble-free operation of quicklime feeders and slakers, keep dust and vapor removal equipment operating properly. Spend at least 10 or 15 minutes of every shift checking and cleaning dust removal device to prevent steam and moisture from surrounding feeding mechanism, causing rusting, premature slaking of lime, and formation of pastes.

b. *Remove Clinkers.* Use hoe or similar tool to remove rocks, clinkers, and heavy sediment from slaker. Keep hoe away from agitators or mixers. In most limes, few clinkers will be found, especially when agitation is of highspeed type.

c. *Clean Slaker.* When slaker is out of service, but not necessarily drained, thoroughly clean each compartment of rocks and accumulated sediment. Carefully clean slaker externally. Use rag oiled with ordinary lubricating oil to wipe off lime feeders and slakers. Quicklime splashing on oily surface does not stick to or affect paint and can be wiped off easily.

(1) *Check Shutoff, Float, and Check Valves for Leakage.* Operate all valves not in daily use. Check for leakage. Clean where necessary.

(2) *Clean Vapor-Removal Apparatus.* When slaker is out of service, remove inspection or cleanout covers and plates for vapor-removal apparatus and clean interior. With covers removed, check mechanism for proper functioning.

d. *Inspect Agitators, Stirrers, and Heat-Exchanger Coils.* Once a month or oftener, drain slakers to examine and clean any copper heat-exchanger coils and check impellers or stirring mechanism. Replace worn impellers where necessary. High-speed impellers on some slakers can be checked without draining slaker by unbolting agitator-shaft bearings and withdrawing agitator shaft with its impeller. If baffles in front of copper heat-exchanger coil are wearing through or improperly shield copper coils from abrasive action of impeller jet, replace them or change position as necessary.

e. *Overhaul.* Drain and thoroughly clean slaker and dust-evacuator system. Examine sides and bottom of slaker for wear and make necessary repairs. Paint slaker lids, inside top edges, and outside to prevent corrosion. Examine and test heat exchanger for leaks and objectionable incrustation. Clean thermometers and test against laboratory thermometer. Disassemble bearings, clean, and lubricate. Clean and repair screen, piping, floats, valves, and controls. Clean vapor removal equipment, paint if necessary.

f. *Clean and Inspect for Corrosion and Mechanical and Electrical Defects.* Clean carefully. Inspect for leaks, wiring defects, metal deterioration, and vibration. Tighten all nuts and bolts. Check belt tension. Paint where necessary to prevent corrosion.

(Chemistry/Bacteriology #22)
Competency Sheet

COMPETENCY: Determine the rate of feed for each chemical solution feeder in your treatment plant and prepare a solution of uniform strength for use in each solution feeder in your plant.*

Objective: Identify the solution feeders in your plant and the chemical being fed by each one.

Objective: Determine the appropriate rate of feed for each chemical solution feeder in your treatment plant.

Objective: Accurately prepare a solution of uniform strength for use in each solution feeder in your plant.

RESOURCES:

(a) Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XIII. -*Gold Appendix*- See section titled "Solution Feeders".

(b) "Solution Feeders" Resource Sheet.

(c) Manufacturer's operators' manual for each solution feeder in your plant.

LEARNING ACTIVITIES:

(a) Read the Resources listed above to find information about chemical solution feeders.

(b) Take a "tour" of your treatment plant. Identify each solution feeder and the chemical it is feeding. Obtain and review the manufacturer's operators' manual for each solution feeder.

(c) Under supervision of an experienced operator, practice preparing solutions of uniform strength for use in each solution feeder in your treatment plant. Be sure to follow the manufacturer's instructions for each solution feeder.

EVALUATION ACTIVITY: While your evaluator observes, accurately prepare the solutions necessary for each solution feeder and add them to the correct solution feeders in your treatment plant. Adjust each feeder to deliver the chemical solution at an appropriate rate.

*If there are no chemical solution feeders in your treatment plant, make arrangements with your evaluator to visit a plant with solution feeders.

"SOLUTION FEEDERS"*

Pump Type Solution Feeders

a. *Check Operation.* Inspect sight-feed indicators to insure that solution is being fed.

(1) See that automatic controls start and stop properly.

(2) Make sure there is enough prepared solution.

(3) Check for leaks in piping joints and packing glands.

b. *Clean Feeder.* Remove and clean glass and plastic parts of sight-feed indicator. Remove white coating caused by hard water in the hypochlorite solution by soaking in 5 percent solution of muriatic acid.

c. *Clean and Flush Solution Tank:*

(1) Clean suction strainer and hose.

(2) Disassemble, inspect, and clean solution diaphragm chambers and check valves.

(3) Clean screens and strainers on water-operated feeders.

(4) Exercise all shutoff valves and length-of-stroke controls.

d. *Overhaul Feeder:*

(1) Replace worn parts and packing.

(2) Clean screen in water meter.

(3) Touch up and paint all metal parts.

Proportional Type Solution Feeders

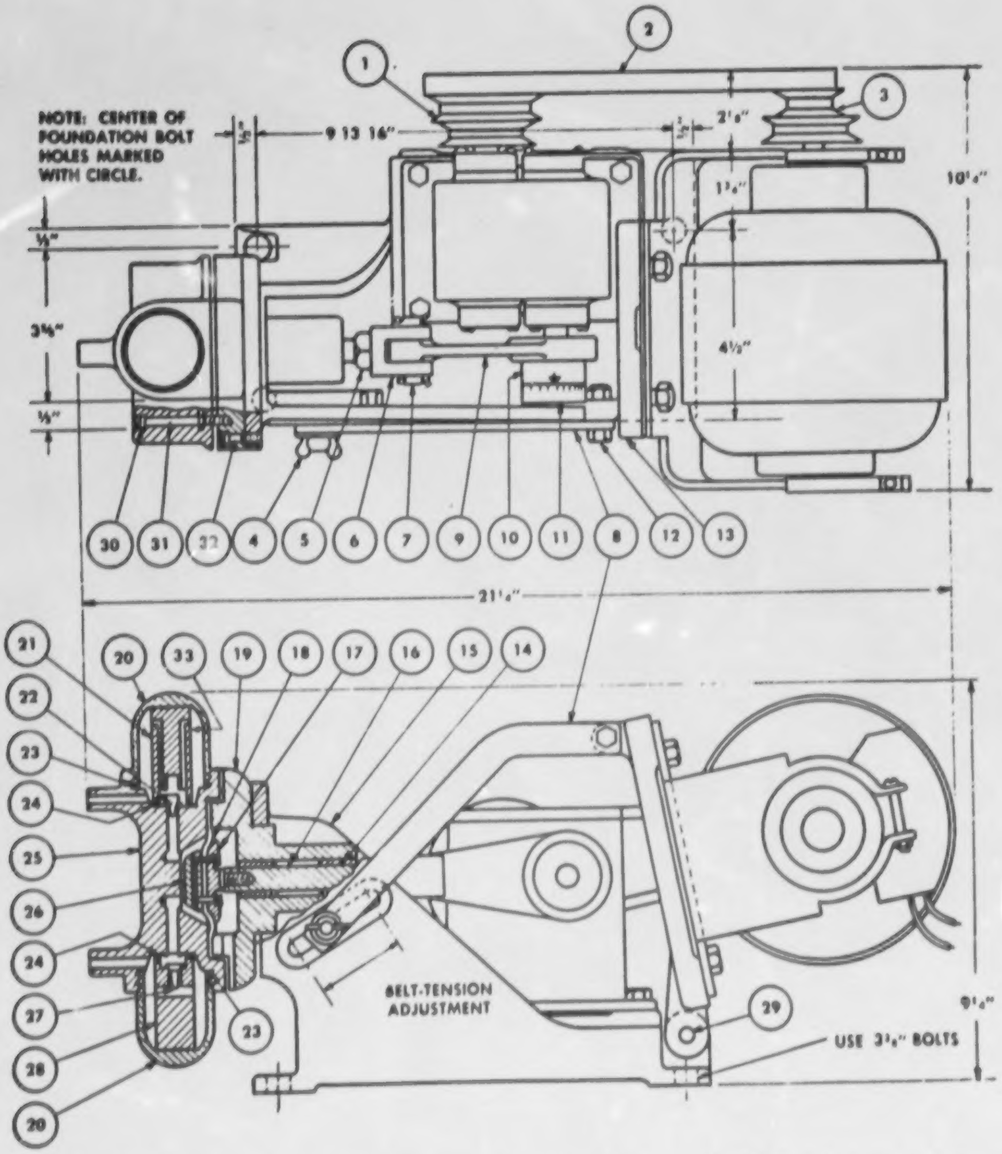
a. *Inspect Sight-feed Indicators.* Make sure solution is being fed. Check operation of unit.

(1) Observe changes in rate of hypochlorite feed under varying rates of flow of water being chlorinated.

(2) Check sufficiency of prepared solution or chemical in feeder.

*Source: Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual #85-13, June 15, 1968, pp. 414-418.

PUMP TYPE SOLUTION FEEDER


NOTE: CENTER OF FOUNDATION BOLT HOLES MARKED WITH CIRCLE.
9 13/16"
2⅛"
1¾"
10¼"
½"
3⅝"
4½"
½"
21¼"
BELT-TENSION ADJUSTMENT
9¼"
USE 3⅜" BOLTS
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30
31
32
4
5
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1. Three-step V-belt pulley (½" bore)
2. V-belt
3. Three-step V-belt pulley (½" bore)
4. 5/16" 18 x 1" long, square-head bolt and wing nut
5. Clevis nut
6. Clevis
7. Clevis pin
8. Motor adjusting bar
9. Eccentric link
10. Outer eccentric (complete)
11. Inner eccentric (complete)
12. 5/16" 18 x 1" long, round-head screw and nut
13. Motor base
14. Push-rod bushing
15. Support frame
16. Push rod
17. Reagent-diaphragm push plate
18. Reagent diaphragm
19. Diaphragm base
20. Check-valve cap
21. Upper check-valve socket
22. Weighted valve tit
23. Check-valve-cap gasket
24. Check-valve-socket washer
25. Reagent-diaghragm
26. Reagent-diaphragm clamp
27. Check-valve tit
28. Lower-check-valve socket
29. Motor-base pivot pin
30. Diaphragm cap-screw washer
31. 10-32 x 1¾" long, round-head machine screw
32. No. 10-32 x ¾" long, filister-head machine screw
33. Sight-feeder intensifier

(3) Check for leaks in piping joints and gaskets.

b. *Clean Feeder.* Remove and clean sight-feed indicator glass and plastic parts.

(1) Clean and flush the solution and chemical dissolving tanks.

(2) Clean screens and strainers on water lines.

(3) Operate all shutoff and rate-control valves.

c. *Inspect Interior.* Look for breaks in special linings. Check for corrosion. Where lining has deteriorated, clean to bare metal and reline with proper material, following manufacturer's instructions.

d. *Overhaul Feeder:*

(1) Remove, clean and inspect all bags or diaphragms.

(2) Remove, clean and inspect all orifice plates, needle valves, etc.

(3) Touch up all metal parts with paint; repaint with enamel if necessary.

BACTERIOLOGICAL SAMPLING AND EXAMINATION*

The collection, transportation, and examination of samples requires considerable effort and expense on the part of the water works and the State Laboratory. If properly collected from representative sampling points, the results of the tests are very important, since they reveal the quality of water the people are being served. The reports are filed and constitute a permanent record relative to the bacteriological quality of the water. These records can be of benefit to the water works officials or the town. When a sample indicates the water is unsatisfactory, the Health Department requests that check samples be collected and submitted. Many times the follow-up check samples must be submitted merely because the water works operator was careless or did not follow instructions. This is both costly and time consuming and many times could be avoided by following instructions closely. Water samples over 30 hours old are not acceptable. This stresses the importance of properly dating and showing time of collection on the identification card.

The laboratory will not test samples which are submitted in bottles other than those furnished by the laboratory for this purpose. This rule is followed as it is necessary to use specially prepared bottles. Since the sample bottle leaves the laboratory clean and sterile, it must be assumed that the water sample is not changed when collected in this bottle. If it is found to be contaminated, it must be assumed that either the water was "bad" or it was improperly collected. The laboratory uses a standard bottle. It is washed, a very small amount of sodium thiosulfate added, capped, and then heat sterilized in an autoclave (which is similar to a large pressure cooker). The sodium thiosulfate is added to react with the chlorine which, in most cases, will be in the public water supply sample. If the chlorine were not neutralized and allowed to remain in the sample, it would kill all bacteria which might have been present when collected. This would cause a false report to be made, and the water works operator would not know the true quality of the water.

The samples are tested for chlorine in the laboratory. When chlorine is found in the sample, it usually indicates the bottle was rinsed or that chlorine was added to try to assure a good report. DO NOT RINSE THE SAMPLE BOTTLE WHEN COLLECTING SAMPLES AND DO NOT ADD CHLORINE TO THE SAMPLE. The purpose of collecting samples is to find "bad" water if it exists in the system. Due to several reasons there may be sections of the distribution system where the water is becoming contaminated, and these should be found if possible. Therefore, samples should be collected from various representative spigots over the entire system. Cross-connections, back siphonage, leaks and other unknown conditions may cause contamination in any section of any system. Part of the job is to seek, find and correct

*Source: Water Works Operators Manual, Alabama Dept. of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter V.

deficiencies. Perfect samples time after time, for long periods of time, should serve to give rise to such questions as: Is this water always perfect? Is this distribution system perfect? Is the technique of collecting samples perfect? Who is being fooled? What is wrong; all samples are always good?

Sampling Procedure

Step by step instructions necessary for the proper collection of water samples are given below:

Seek a sampling station which should give representative results of water in the area. If possible, collect all samples from a yard tap or faucet. Chrome plumbing fixtures, usually found inside, may be damaged or made unsightly by properly flaming the fixture to sterilize it. This is the reason for recommending outside spigots. Avoid taking samples from garages, filling stations, kitchens or other places where grease may collect on the tap. Samples should not be collected from fire plugs. Fire plugs have weep holes and large barrels which may be easily contaminated and are very difficult to disinfect. The mouth of the fire plug cannot be disinfected easily; nor can a bottle be properly filled. After the spigot is selected:

(1) Let the water run for several minutes to insure fresh water from the system. Shut water off.
(2) Then flame the faucet in order to insure the destruction of any bacteria which may be clinging to the faucet. If available, a blowtorch can be used for this purpose. There are small "easy to light" (L.P. or butane gas) torches available which are excellent for this purpose. The flame should be directed against the inside edge of the faucet as much as possible. A good torch can be made by tightly rolling a piece of cotton or a rag, wrapping it with a stiff wire, leaving about a foot of the wire extended for a handle. Soak this in alcohol. This will light easily, makes a good flame, and can be lit and put out several times before fuel must be added. A rolled up newspaper can be used. Do not overheat the fixture.
(3) Open the faucet and let the water run in a small stream about the size of a pencil.

(4) While the tap is running, remove the lid from the sample bottle. Do not touch the rim of the bottle or the inside of the lid.

(5) Hold the lid in one hand, and quickly fill the sample bottle to approximately three-fourths full. You must not fill the bottle to brimming over; the State Hygienic Lab will not examine such a sample, since an air space is required between the lid and the water surface. When the bottle contains enough water (over 100 cc.), remove it from under the tap and immediately replace the cap. Turn it down firmly to insure a good seal between the bottle and cap.

(6) Fill out the identification form completely, and keep it with the sample for mailing.

(7) SHIPPING. The samples should be collected as near to shipping time as possible since the sample must reach the testing laboratory within 30 hours after collection. If the case and a supply of clean, sterile bottles are to be returned by mail, postage must be included. The shipping container will be returned at the water works expense. At the present time, the samples do not need to be iced during shipment as previously required. Prompt shipment cannot be overstressed, since the sample must be at the laboratory within 30 hours after it is collected. It will probably require checking into shipping, mail, and bus schedules in order to meet the 30-hour deadline.

The above directions must be followed explicitly in order to get reliable results.

After the sample arrives at the laboratory, the bacteriological examination is made and the report of examination is sent to the water system.

DO NOT TAKE SAMPLES FROM FIRE HYDRANTS!

Bacteriological Examination

When the water sample arrives at the laboratory, the bacteriological examination is accomplished by the membrane filter method.

The membrane filter procedure has now gained official acceptance from the U.S.P.H.S. and the American Water Works Association. It was first described as a full alternate procedure in the Twelfth Edition of Standard Methods for the Examination of Water and Wastewater (1965).

All operators should be familiar with the principles of the bacteriological tests as performed by the laboratory. This knowledge should help him to interpret the results reported to him by the State Health Department. In order that the operator may become acquainted with the principles, a brief discussion is given below on both the membrane filter method and the multiple tube fermentation method. The multiple tube method is the procedure that was universally used prior to the development of the membrane filter method and is still an approved method for bacteriological examination today.

The most recent edition of "Standard Methods" should be used as reference in setting up a laboratory and making bacteriological tests.

Membrane Filters

Membrane or MF Millipore Filters have come into use in many fields of bacteriology and have supplemented or replaced some of the older methods of isolating and identifying bacteria. These filters are thin membranes made of bonded cellulose ester sheet material approximately 150 microns thick (0.006 inch). The structure of the filters is such that they have uniform pores which compose 80 to 85 percent of the membrane surface. By varying the size of the cellulose ester particles, the size of the pores can be controlled and the membranes can be made to be used for many different purposes. The filters are used for bacteriological examinations by forcing liquid samples through the membrane with pressure or suction. The pores of the membrane will permit water to pass through; however, they are of such size that bacteria cannot pass but are retained on the surface of the filter. For example, the Type HA filters which are used by the Health Department Laboratories have pores which are 0.45 microns (20 millionths of an inch) in size; and they will retain nearly all known non-viral microorganisms.

Extensive work with the membrane filter on the examination of water samples for coliform bacteria has been done; and it is now possible to filter a sample of water through a membrane filter, then grow out coliform bacteria on the surface of the filter and get a direct or actual coliform count on a definite volume of water.

The membrane filter method offers several advantages over the multiple tube fermentation procedure. These include:

(1) Time required for the test for coliform bacteria is reduced from 48 or 96 hours to 20 hours.
(2) A larger and more representative sample of water can be subjected to the test, thus increasing accuracy.
(3) The test can be made in the field as well as in the laboratory.
(4) Direct reading of the bacteria count is made rather than arriving at the "most probable number" of bacteria by a statistical estimate.
(5) Specific pathogens can be isolated with a relative ease not heretofore possible (in part, due to the enormously large sample size permitted).
(6) The simplicity of use of the membrane filter and of the analysis of results obtained by its use has permitted a sharp reduction in the technician hours required to establish adequate water controls, while at the same time assuring greater reliability of results.

Perhaps the most outstanding feature of the procedure is the availability of results in a very much shorter period of time which permits prompt initiation of corrective measures where indicated and places this control on a day to day basis. Incidentally, this very feature permits early release to service of newly installed pipelines. Previous methods often served only as a historical record of seasonal and other large variations in water conditions.

Using the membrane filter, the laboratory tests of a sample consist of three simple steps:

1. Filtration
2. Incubation
3. Observation of results

Water samples are collected in a sterile container from the water source. A portion, usually 100 ml., of each sample is then filtered through the membrane filter held in a standardized filter holder. The filters and apparatus are easily sterilized by any of several simple procedures; however, care must be exercised in their use. While the water passes easily through the filter, bacteria are unable to do so and are deposited on its surface. A water sample of the size indicated would pass through a filter in about one minute.

The filter bearing the bacteria on its surface is then placed on a nutrient pad (food for bacteria) in a plastic petri dish. The nutrient solution must be in accordance with "Standard Methods" and can be bought already prepared. The petri dish holding the filter and the nutrient solution is then placed in an incubator at approximately body temperature (98 degrees F.) for 20 hours.

During the incubation period, a visible bacterial colony will grow wherever a single organism was caught by the filter. After incubation, the filter disc is removed from the dish and allowed to dry for one hour on absorbent paper. The coliform colonies are then counted.

The coliform colonies are easily distinguished from other bacteria that may be presented because they exhibit a characteristic and unique golden, metallic sheen. The MF-Endo Broth helps to bring out this difference and it also suppresses the growth of most non-coliform bacteria. Through its use, the coliforms are easily identified.

The number of "sheen" colonies on the filter disc reveals the number of coliform bacteria in the volume of the filtered sample. The laboratory reports the results in terms of the number of coliforms per 100 ml. of water. If 200 ml. of water are filtered and two "sheen" colonies develop, the report shows one coliform per 100 ml.

Bacterial examination of water in the field has been made practicable by the development of a field testing kit. In this kit, the membrane filter disc and absorbent pad are assembled in a two-part sterile plastic device which serves as a filter holder and culturing dish. This assembly is called a "field monitor." To use the field monitoring set a sample of water is forced through a membrane that has been properly prepared and enclosed in a plastic holder to protect it. The filter is then saturated with medium or food for growing bacteria and the two openings in the monitor are plugged. It is then incubated in an inverted position in a portable incubator that is available, or the filter can be mailed to the laboratory for incubation and counting of coliform colonies. The use of these sets may make present methods of shipping water samples into the laboratory unnecessary.

Plate Count and Multiple Tube Method

As stated previously, prior to the membrane filter method, the multiple tube method was used for bacteriological examination of water. In conjunction with this test, a total plate count was also conducted. These tests are discussed here since they are still considered standard tests in the water works field.

When a water sample arrives at the laboratory, two tests--the plate count and the coliform test by the multiple tube method--are made and reported to the operator. The coliform test actually consists of two steps known as the presumptive test and the confirmed test. Under certain conditions, it is necessary to go one step further and make a completed test, however, this step is not always necessary. To make the tests, small portions of the water sample are used in accordance with the following procedures.

The Plate Count

The plate count is a test made by the laboratory to determine the total number of bacteria present in the sample. This test does not differentiate between the many different types of bacteria and is thought of as giving index to the general "housekeeping" practices. A "high" count indicates that some type of contamination is present and is undesirable.

The test is made by placing a portion of agar in a petri dish. Agar is a type of plant food that favors the growth of most bacteria. A one cc portion of the water sample is placed in the petri dish along with the agar. It is then placed in an incubator with the temperature at 37 degrees C. or 98.6 degrees F., which is body temperature. After 24 hours, the plate is removed, examined, and the colonies in and on the agar are counted and recorded on the report form as "Bacteria per cc at 37 degrees C."

It is assumed that each single bacteria in the agar starts to grow and due to the short generation time, it develops so that a spot or colony visible to the eye can be seen on the plate. It is these colonies that are counted. Each colony is considered to represent an original bacterium.

The Multiple Tube Test

The following tests are to determine if coliform organisms are present and are much more significant than the plate count. The presence of these organisms may be an indication that harmful bacteria are entering the supply.

The presumptive test: Like some seed, coliform organisms are so small that they cannot be seen with the naked eye. Again, like seed, if they are to grow, they must be planted where food is available and other conditions are favorable. Since a good food for coliform and a few other organisms is Lactose Broth, a portion of the water sample, which may contain the seed or bacteria, is planted in test tubes containing this food. The Lactose Broth also provides the moisture needed for growth. The test tubes are placed in the incubator which provides the best temperature, 98.6 degrees F.

If food, moisture, and proper temperature are provided, the organisms, if present, will grow if given sufficient time. If coliform organisms or the few other types which will grow in Lactose are not present in the water, then naturally nothing happens, as there is nothing to grow. However, if organisms are present, they will grow under these ideal conditions and will ferment the Lactose Broth and produce gas in it. The gas indicates the presence of the organisms.

After 24 hours in the incubator, the tubes are examined for gas. If no gas has formed, they are given an additional 24 hours time. If, after 48 hours, gas has not formed, then no organisms were present and the report reads "Absent" and the water is considered safe for drinking purposes.

Now, if organisms that reproduce in Lactose Broth are present in the sample, gas will be produced in any or all of the tubes within 48 hours. Because of the presence of gas, it is known that some type organism is present; and it is presumed that the organisms are coliform. That is the reason the test is called the "presumptive test", as it is merely presumed coliform organisms are present if gas is produced in Lactose Broth.

The confirmed test: The test would have been completed if the coliform organisms were the only ones which grow in Lactose Broth. Unfortunately, other organisms will ferment the broth and produce gas. Thus, it becomes necessary to transfer a small amount of the contents of the Lactose Broth tubes showing gas into another, more selective food called Brilliant Green Bile. Coliform organisms are the only organisms which are able to grow in this food. Since these organisms are found in the colon and intestines, it might be said that they are more or less at home in this food containing bile.

The tubes of Brilliant Green Bile containing some of the organisms from the Lactose Broth are then placed in the incubator. At the end of 24 and 48 hours, they are examined for gas. If gas is present, the confirmed test is positive and the results are reported as "Present". The water contained coliforms and must, therefore, be questionable as a safe drinking water. If no gas is present, the confirmed test is "negative" and the results are reported as "Absent", and the water is considered safe.

The presence of the Coli-aerogenes group of bacteria in the above tests does not definitely mean that harmful bacteria are present. Colirom bacteria are normally present in great numbers in the human intestine and, except in unusual circumstances, are not harmful to humans. When present in a water sample, however, they do indicate the presence of fecal contamination and the possibility that harmful (pathogenic) organisms, such as typhoid fever germs, may be present. Therefore, the tests are not measures of actual disease-producing organisms, but rather are indicators of the possibility that such are present.

"SOFTENING HARD WATER"*

Softening (Lime-Soda)

a. *General.* Although hard water may be potable, it is objectionable because it deposits scale in pipes and boilers, leaves stains on glassware and kitchen utensils, and requires excessive amounts of soap to form lather. The hardness of water can be judged from the amount of soap used to form lather. Hardness results from the presence of calcium and magnesium compounds in solution in the water. Water containing over 200 ppm of hardness expressed as $CaCO_3$ is considered very hard, 100-200 ppm as hard, 50-100 ppm as moderately hard, and below 50 ppm as soft. Softening of hard water is achieved by changing the calcium and magnesium compounds from a soluble to an insoluble form and then removing the insoluble compounds by sedimentation and filtration. Softening also is accomplished by the zeolite or base exchange process. The most common process of softening by precipitation is known as the lime-soda process. Lime-soda process plants are essentially the same as water filtration plants. Lime and soda ash are added to raw water, and the softening reaction occurs during mixing and flocculation; the precipitated calcium and magnesium are removed during sedimentation. An additional step, called recarbonation, is frequently provided just prior to filtration. If raw water has a high turbidity, the turbidity is partially removed by plain sedimentation prior to the softening process.

b. *Hardness Producing Chemicals.* Hardness is generally classified as carbonate or noncarbonate hardness, depending on the chemical compounds that cause it.

(1) *Carbonate Hardness.* Carbonate hardness, formerly referred to as temporary hardness, is caused by calcium or magnesium bicarbonate in the water. They form when carbon dioxide combines with water to form a mild acid which in turn combines with calcium and magnesium carbonate in the surrounding soil. The carbonate is insoluble; the bicarbonate which is formed dissolves in the water.

(2) *Noncarbonate Hardness.* Noncarbonate hardness, formerly referred to as permanent hardness, is caused by:

(a) Calcium and magnesium sulfate.

(b) Calcium and magnesium chloride.

(c) Calcium and magnesium nitrate.

Note: Iron and aluminum compounds also produce hardness, but usually are present in such small quantities that they are not generally associated with hardness. Water softening processes remove iron as well as calcium and magnesium salts.

*Source: Maintenance and Operation of Water Plants and Systems, Dept. of the Air Force, Air Force Manual #85-13, June 15, 1968, pp. 162-68, pp. 207-208.

c. *Lime-Soda Process:*

(1) *Softening Reactions.* CaO (calcium oxide) is added to water forming calcium hydroxide:

CaO	+	H_2O	=	$Ca\ (OH)_2$
Calcium Oxide		Water		Calcium Hydroxide

(a) When lime is added to water containing carbonate (temporary) hardness, the calcium hydroxide reacts with calcium and magnesium bicarbonates to form less soluble carbonates and hydroxides that precipitate out, leaving the water free of the substances causing carbonate hardness:

$Ca(HCO_3)_2$	+	$Ca(OH)_2$	=	$2CaCO_3$	+	$2H_2O$
calcium bicarbonate		calcium hydroxide		calcium carbonate precipitate		water

$Mg(HCO_3)_2$	+	$2Ca(OH)_2$	=	$Mg(OH)_2$	+	$2CaCO_3$	+	$2H_2O$
magnesium bicarbonate		calcium hydroxide		magnesium hydroxide precipitate		calcium carbonate precipitate		water

A similar chemical reaction occurs when slaked or hydrated lime is used.

(b) In the lime-soda process, noncarbonate (permanent) hardness is removed from the water by use of soda ash, Na_2CO_3 (sodium carbonate). When soda ash is added to water containing noncarbonate hardness the following chemical reactions take place:

$CaSO_4$	+	Na_2CO_3	=	$CaCO_3$	+	Na_2SO_4
calcium sulfate		sodium carbonate		calcium carbonate precipitate		sodium sulfate
$CaCl_2$	+	Na_2CO_3	=	$CaCO_3$	+	$2NaCl$
calcium chloride		sodium carbonate		calcium carbonate precipitate		sodium chloride
$Ca(NO_3)_2$	+	Na_2CO_3	=	$CaCO_3$	+	$2NaNO_3$
calcium nitrate		sodium carbonate		calcium carbonate precipitate		sodium nitrate

(c) Soda ash reacts with sulfates, chlorides, and nitrates of calcium to form insoluble precipitates. If the water contains both carbonate and noncarbonate hardness, as is usually the case, both lime and soda are used.

(2) *Determination of Dosage:*

(a) Quantitative analysis of the water is necessary for determination of required dosage of lime and soda ash.

(b) It is well to check the calculations for lime and soda ash requirements by actually adding the calculated dosages to small measured quantities of the water to be treated. The following procedure will be found convenient.

1. Prepare a standard lime suspension by adding 10.0 grams of lime to 1 liter of boiled and cooled distilled water. If quicklime is used, shake it with just enough water to make a boiling solution, then dilute to 1 liter. A standard solution of soda ash is prepared in the same way. Adding 1 ml of the standard solution to 1 liter of water gives a dosage of 10 ppm.

2. Place 1 liter of water to be treated in a beaker or jar and add calculated quantities of lime and soda ash needed to soften it. Stir gently for 30 minutes and allow to settle until liquid becomes fairly clear, usually 10 to 15 minutes.

3. Siphon off about half the clear solution, warm to about 120 degrees F (49 degrees C), and filter through filter paper.

4. Analyze filtrate for alkalinity in accordance with standard laboratory test procedure.

5. Use tables at the end of this Appendix, pages 8 and 9, to determine extent of softening.

(c) In the operation of a softening plant, it is customary to use lime in excess of computed needs to the extent of 10 to 50 ppm. This is then removed by recarbonation. After recarbonation and filtration, the water should have an alkalinity of approximately 35 ppm.

(d) The quantity of noncarbonate hardness allowed to remain in the water is regulated by decreasing soda-ash feed in accordance with the total hardness desired. If the total hardness in the softened water is to be 85 ppm and the alkalinity 35 ppm, it is then necessary to remove all but 50 ppm of noncarbonate hardness.

d. *Chemical Feeding.* Continuous softening requires application of chemicals in proportion to the rate of flow of water through the treating units. If the flow is constant over definite periods of time, the chemicals can be fed through adjustable constant-feed machines. If the flow is variable, proportioning devices in which chemicals are fed in solution or suspension are used.

(1) *Soda Ash.* Soda ash dissolves easily in water, forming a clear stable solution. After preparation, the solution needs no further agitation and can be added directly to the water through a proportioning device.

(2) *Lime.* Lime, being only slightly soluble in water, forms a suspension known as milk of lime. Because suspended matter tends to settle, continuous agitation is necessary to keep the mixture uniform. Therefore, devices for feeding lime usually have mechanical agitators. Flow through feed lines must be fast enough to prevent lime from settling and clogging the lines. In long lines this requires large quantities of dilution water. Introducing lime suspension through a trapless pipe which terminates above the raw water eliminates clogging. Long lines should be avoided or should be easily replaceable. They should have pressure connections for flushing. Rubber hose or open troughs are frequently used for this purpose.

(3) *Lime-Soda Ash.* Some plants feed lime and soda ash through one proportioning device. In some cases, equipment like that for feeding milk of lime is used. Chemicals are mixed in the proper proportions with water and introduced as a thin slurry. This method is satisfactory where the composition of raw water is constant or does not change rapidly.

e. *Supplementary Softening Procedures.* The lime-soda softening process often results in formation of soluble complex basic carbonates of magnesium and high alkalinity of softened water. Methods of overcoming this limitation are discussed below.

(1) *Excess-Lime Treatment.* In the excess-lime treatment, 35 to 50 ppm excess lime or more is added to precipitate the magnesium almost completely. Excess calcium is then precipitated with soda ash. Soda-ash dosage is the amount required to combine with the noncarbonate hardness of water and excess lime. If the water is recarbonated after filtration to a pH of 8.0, it will be palatable and will have a hardness of about 30 ppm.

(2) *Split Treatment.* In water treatment plants which were not designed for softening, inadequacy of softening facilities can be overcome by split treatment. In this process, as large a portion of the raw water as possible is over-treated to obtain maximum reduction of hardness. The excess is then neutralized with raw water. Split treatment produces a greater reduction in hardness of a given quantity of raw water than does addition of the same quantity of chemicals to the total supply.

(3) *Excess Lime and Recarbonation.* Treatment with excess lime and recarbonation consists of the excess-lime treatment and use of carbon dioxide to neutralize the excess lime. Carbonate hardness can be reduced by this method approximately to the solubility limit, 20 to 35 ppm hardness (depending on temperature). Calcium carbonate is least soluble at pH of 9.3 to 9.7.

(4) *Addition of Aluminum Compounds.* Aluminum compounds coagulate finely divided precipitates, thereby facilitating their removal. Without aluminum treatment, longer reaction periods are required for precipitates to coagulate and settle. Addition of aluminum compounds converts soluble magnesium salts to insoluble magnesium aluminates which can be removed by settling or filtration.

(5) *Excess Lime, Recarbonation, and Secondary Flocculation.* Treatment with excess lime, recarbonation, and secondary flocculation is carried out as follows:

(a) Add excess lime and correct amount of soda ash to combine with noncarbonate hardness; mix and allow to settle.

(b) Neutralize excess lime with carbon dioxide gas.

(c) Mix with sludge produced by the water softening process or with a coagulant and allow to settle.

(d) Add carbon dioxide gas again and filter. This process decreases load on filters, lessens incrustation of filter sands, and produces a more nearly balanced and therefore more stable filtered water. It is especially adapted to treating and softening highly turbid water.

f. *High Rate Softening Devices.* Newer developments in softening include high rate equipment such as the *Precipitator, Accelator, Hydro-Treator,* and *Spiractor.* The *Precipitator, Accelator* and *Hydro-Treator* are also used as combined flocculation and sedimentation units without softening. When these units are operated before filtration to treat surface water with low suspended solids and low alkalinity, it may be necessary to add lime or clay to add weight and prevent rising floc. Wide fluctuations in flow, solids content, and temperature also cause operating problems which can best be corrected individually.

(1) *Accelator, Precipitator, and Hydro-Treator.* The Accelator (Figure 1), the Precipitator, and Hydro-Treator (Figure 2) operate on essentially the same principles. Precipitates formed are kept in motion by a combination of mechanical agitation and hydraulic flow. Velocity of flow of water through the system is controlled to keep precipitates in suspension at a level where water passes through them. The accumulated precipitate is called the sludge blanket. When operating properly, water above the sludge blanket and flowing over the weirs is clear. Operation depends on balancing the lift of particles by the velocity of upward flowing water against the pull of gravity. Thus, if the velocity of water is high and then is gradually decreased, a point is reached at which the particles are too heavy to be supported by the velocity of the water and remain behind, thus separating from the water. Continuous treatment builds up the sludge blanket which is drawn off as required. Operation of the equipment is covered in detail in manufacturers' instruction manuals. Problems not covered by them should be referred to higher technical authority.

(2) *Spiractor.* The spiractor (Figure 3) consists of an inverted conical tank in which the lime-soda softening reactions take place in the presence of a suspended bed of granular calcium carbonate which acts as a catalyst. In operation, the tank is slightly more than half filled with 0.1 to 0.2 mm granules. Hard water and chemicals enter the bottom of the cone tangentially and close to each other so that they mix immediately as the treated

Figure 1

Suspended Solids Clarifier (Accelator)

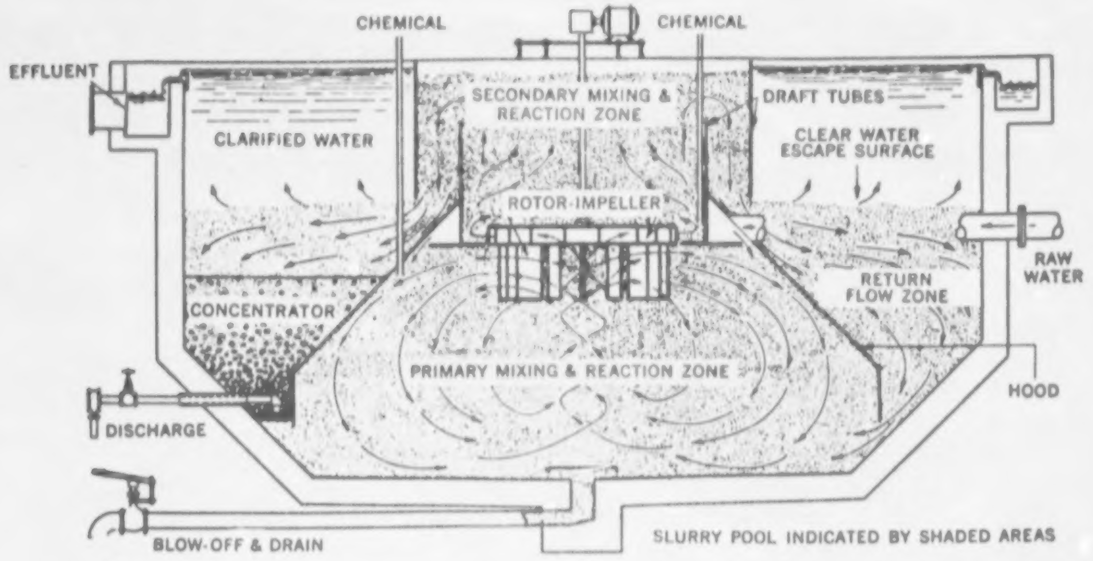

CHEMICAL
CHEMICAL
EFFLUENT
SECONDARY MIXING & REACTION ZONE
DRAFT TUBES
CLARIFIED WATER
CLEAR WATER ESCAPE SURFACE
ROTOR-IMPELLER
RAW WATER
RETURN FLOW ZONE
CONCENTRATOR
PRIMARY MIXING & REACTION ZONE
HOOD
DISCHARGE
BLOW-OFF & DRAIN
SLURRY POOL INDICATED BY SHADED AREAS

Figure 2

Suspended Solids Clarifier

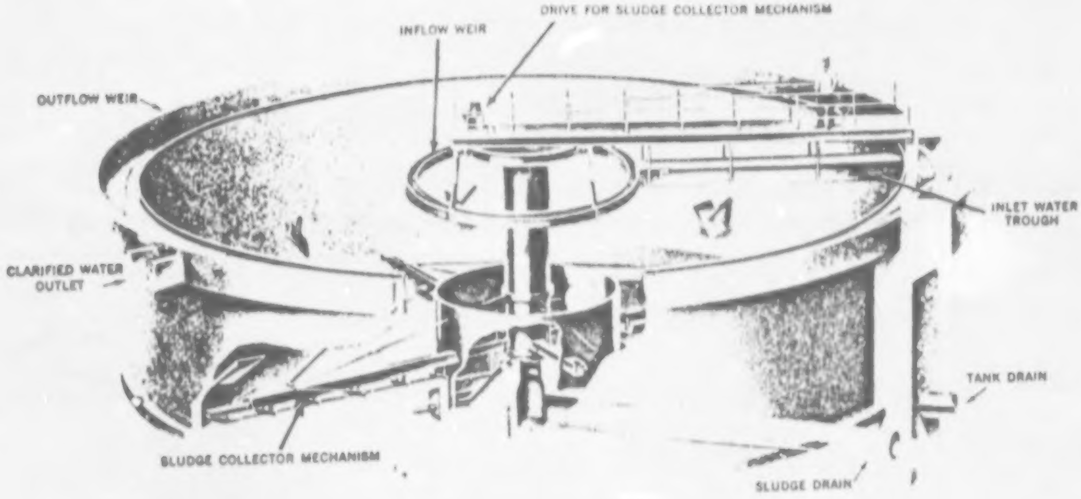
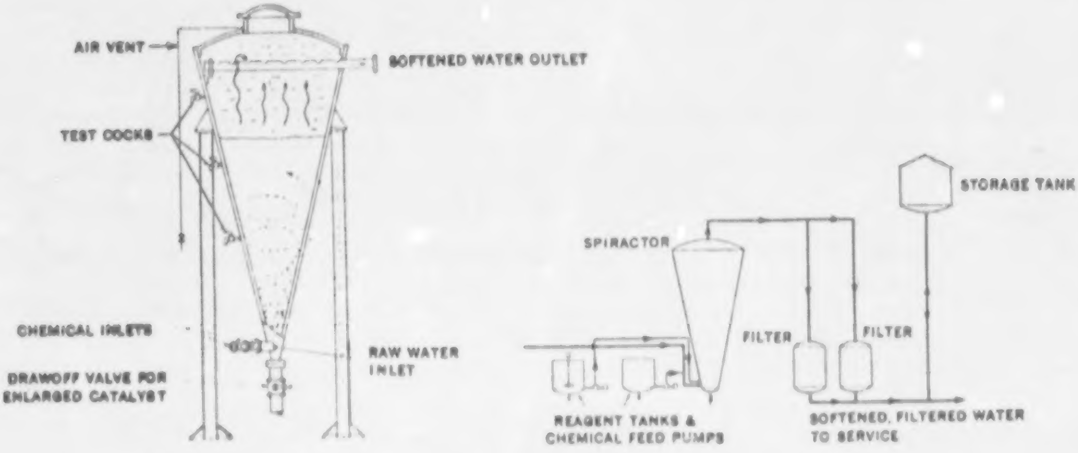

DRIVE FOR SLUDGE COLLECTOR MECHANISM
INFLOW WEIR
OUTFLOW WEIR
INLET WATER TROUGH
CLARIFIED WATER OUTLET
TANK DRAIN
SLUDGE COLLECTOR MECHANISM
SLUDGE DRAIN

Figure 3

Contact Solids Softener (Spiractor)


AIR VENT
SOFTENED WATER OUTLET
TEST COCKS
CHEMICAL INLETS
RAW WATER INLET
DRAWOFF VALVE FOR ENLARGED CATALYST
STORAGE TANK
SPIRACTOR
FILTER
FILTER
REAGENT TANKS & CHEMICAL FEED PUMPS
SOFTENED, FILTERED WATER TO SERVICE

SECTION OF SPIRACTOR

SYSTEM

Stream of water rises through the granular bed with a swirling motion. The upward velocity keeps the granular material in suspension. As the water rises, velocity decreases to a point where material is no longer in suspension. The contact time, 8 to 10 minutes, is enough to complete softening reactions. Softened water is drawn off from the top of the cone. The precipitated hardness particles attach themselves to the calcium carbonate granules during the process, increasing the bulk of granules in the unit. However, the level of the unit is kept down to the desired point by withdrawing the largest particles from the bottom. New catalyst material must be added periodically. It can be produced by regrinding and screening the discharged material. The softened water is usually filtered through a sand filter to remove remaining turbidity. Advantages of the equipment are its small size, low installation cost, rapid treatment, and lack of moving parts and pumping equipment. In addition, the precipitated hardness is withdrawn as sand-like granules, which are much easier to dispose of than the watery slime from the normal lime-soda process. The unit is most effective when hardness is predominately calcium; there is less than 17 ppm magnesium hardness (expressed as calcium carbonate), water temperature is about 50° F., and turbidity is less than 5 ppm.

ESTIMATING ALKALINE COMPOUNDS

General. In the lime softening process, the alkaline-compound constituents can be readily found by using the results of alkalinity and total hardness tests and Tables 5, 6, and 7. Table 5 gives alkalinity due to hydroxides, carbonates, and bicarbonates from a comparison of the methyl orange and phenolphthalein titrations. Table 6 gives the soda alkalinity, noncarbonate hardness, and carbonate hardness from a comparison of methyl orange alkalinity and total hardness. In Table 7, caustic alkalinity (hydroxide) as determined from Table 5 is compared with soda alkalinity as determined from Table 6. This comparison gives estimations of concentrations of sodium hydroxide, sodium carbonate, normal carbonates of calcium and magnesium, and calcium hydroxide.

TABLE 5. COMPARISON OF PHENOLPHTHALEIN AND METHYL ORANGE ALKALINITIES

	OH	CO_3	HCO_3
When P = 0	0	0	M
When P < ½M	0	2P	M-2P
When P = ½M	0	2P	0
When P > ½M	2P-M	2(M-P)	0
When P = M	M	0	0

TABLE 6. COMPARISON OF TOTAL HARDNESS AND METHYL ORANGE ALKALINITY

	Na Alk	Non CO_3 Hardness	CO_3 Hardness
When TH < M.....	M-TH	0	TH
When TH = M.....	0	0	TH
When TH > M.....	0	TH-M	M

Key to Tables

Alkalinity to phenolphthalein	P
Alkalinity to methyl orange	M
Total hardness	TH
Caustic alkalinity	OH
Normal carbonate alkalinity	CO_3
Bicarbonate alkalinity	HCO_3
Soda Alkalinity	Na Alk
Non carbonate hardness	Non Co_3 hardness
Carbonate hardness	CO_3 hardness
Sodium hydroxide	NaOH
Sodium carbonate	Na_2CO_3
Normal carbonates of Ca and Mg .	Ca and Mg CO_3
Calcium hydroxide	Ca $(OH)_2$

TABLE 7. COMPARISON OF CAUSTIC AND SODA ALKALINITIES

	NaOH	Na_2CO_3	Ca and $MgCO_3$	$Ca(OH)_2$
When OH<Na Alk..........	OH	Na Alk-Oh	TH	0
When OH=Na Alk..........	OH	0	CO_3	0
When OH>Na Alk..........	Na Alk	0	CO	OH-Na Alk

b. *Use of Tables.* The two examples below illustrate use of the tables.

(1) *Example 1.* Test of a water sample gave the following results:

M=70 ppm.
P=40 ppm.
TH=105 ppm.

(a) P is greater than one-half M, so line 4, table 5 is used to find OH, CO_3, and HCO_3.

OH=2P−M=80−70=10 ppm.
CO_3=2(M−P)=2(70−40)=60 ppm.
HCO_3=0 ppm.

(b) TH is greater than M, so line 3, table 6 is used to find Na Alk, Non CO_3 hardness, and CO_3 hardness.

Na Alk=0 ppm.
Non CO_3 hardness=TH−M=105−70=35 ppm.
CO_3 hardness=M=70 ppm.

(c) OH is greater than Na Alk, so line 3, table 7 is used to find Na OH, Na_2CO_3, Ca and $MgCO_3$, and $Ca(OH)_2$.

NaOH=Na Alk=0 ppm.
Na_2CO_3=0 ppm.
Ca and Mg CO_3=CO_3=60 ppm.
Ca $(OH)_2$=OH−Na Alk=10−0=10 ppm.

(2) *Example 2.* Test of another water sample gave the following results:

M=60 ppm.
P=40 ppm.
TH=35 ppm.

(a) P is greater than one-half M, so line 4, table 5 is used to find OH, CO_3 and HCO_3.

OH=2P−M=80−60=20 ppm.
HCO_3=0 ppm.

(b) TH is less than M, so line 1, Table 6 is used to find Na Alk, Non CO_3 hardness, and CO_3 hardness.

Na Alk=M−TH=60−35=25 ppm.
Non CO_3 hardness=0 ppm.
CO_3 hardness=TH=35 ppm.

(c) OH is less than Na Alk, so line 1, table 7 is used to find NaOH, Na_2CO_3, Ca and $MgCO_3$, and Ca $(OH)_2$.

$Ca(OH)_2$.
NaOH=OH=20 ppm.
Na_2CO_3=Na Alk−OH=25−20=5 ppm.
Ca and $MgCO_3$=TH=35 ppm.
$Ca(OH)_2$=0 ppm.

APPENDIX
GOLD

COAGULATION*

Following the route of the water through the filtration plant, the first point in the plant it reaches after leaving the intake is some type of mixer. It is at this point, usually, that the first chemicals are added. The exception to this is that there are some supplies where chlorine or other taste and odor control measures are introduced at the intake.

In order to understand why any chemicals, other than chlorine, must be added it is necessary to understand certain things about the water. Depending upon the source, the waters to be treated or purified may vary from waters heavily laden with silt and mud to very clear waters. The dispersed substances that may be present in the water exist as suspended matter, colloids or solution. These will be discussed in the following paragraphs.

First, consider a sample of water from a stream at flood stage. It is very turbid and may even contain relatively large, heavy particles such as flakes of mud, pieces of vegetation, and large grains of sand. As long as the stream flows fast and carries these particles along, they would be known as "suspended matter." As soon as the stream reaches a larger body of water, where its velocity is decreased, this heavy, suspended matter drops or settles out. At times of low flows, the water contains suspended matter (smaller and lighter) which will drop out when the water becomes still. Certain small particles may be classed as suspended, even though they are not heavy enough to settle when the water is completely still.

Other material or matter may be in solution. An example of this is sugar dissolved in water. The colloids or matter in the colloidal state may be considered as being the "in-between" sizes; that is, sizes between suspended particles and dissolved substances. Color, as in "tea colored" swamp waters, is mainly present due to colloids or extremely fine particles of matter in suspension although part of it is in true solution. From the above it can be seen that turbidity may be present in any or all of the three states mentioned while most is usually present as colloids and solutions. These particles in water, in addition to being very small, are also electrically charged. Most charges are alike (negative charge) and, hence, repell each other (as do like poles of a magnet) and, therefore, stay dispersed, thus remaining in the colloidal or soluble state.

Turbidity, caused by the coarse, suspended matter, will settle without any treatment, requiring only that the water be stilled enough to allow it to drop out. Its rate of settling can even be determined, as it will generally follow certain laws of physics (Stokes Law) reasonably well. However, even some of the relatively coarse suspended turbidity may be so slow in settling that plain settling becomes impractical. Knowing that some of the coarser materials do not settle well, it is easily understood that the finer suspensions, colloids and solutions, would never settle. Therefore, plain sedimentation is usually an impractical way to obtain a clear water from streams in Alabama; and some means of increasing the settling rate must be used.

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*Source: Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XIII.

The operator may be interested in knowing that the sizes of the particles discussed above, although very small, can be estimated or realized. They are so small that, when they are measured by scientists, they are spoken of in terms of millimicron, which is abbreviated and commonly written as Mu. Each millimicron is one-millionth of a millimeter, and a millimeter is only 0.039 of an inch. Figure 1 may give some idea as to the size of the particles.

FIGURE 1

Approximate Size Range
of

Type Particles	Particle in Millimicrons (Mμ)	Stability	Visibility
Solution (or dissolved material)	Less than 1 Mμ (1/25,000,000 of an inch)	Completely stable	Not visible
Colloids	1 to 500 Mμ	Can remain suspended for years	Can be seen with ultra microscope
Bacteria	200 to 10,000	Suspended	Visible with microscope
Suspended Matter	Greater than 1,000 Mμ	Can be settled	Visible with microscope and larger particles can be seen with naked eye

Just as the naturally coarse, suspended particles settle, so should the very small ones if they could be gathered together in such a way as to form relatively large, heavy flakes or clusters. There is a way of doing this, and it is known as coagulation. This makes it possible to treat turbid waters so that they can be made clear enough to be filtered. Coagulation may be defined as "the process of causing electrical and chemical reactions to occur that will result in the finely divided materials in water to cluster together, forming what is commonly termed as 'floc'." The floc, when properly formed, will settle. Illustrations of the process of coagulation are the preparation of gelatin (Jello) desserts or heat causing the jelling or coagulation of the white of an egg.

Floc, properly formed, will settle quickly. It is formed by adding to the water certain chemicals known as coagulants. The most common of these are aluminum sulfate (alum) and ferric sulfate. When alum is dissolved in the water, it reacts with the alkalinity in the water. Those whose knowledge of chemistry is very limited may do well to think of the term alkalinity as being lime in the water. The addition of lime to water increases the alkalinity. This reaction results in disassociation to form ions or electrically charged particles. These charged (coagulant) particles contact, neutralize and so act upon the colloidal particles that they begin to attract each other (as do opposite poles of the magnet). As this comes to pass, the particles begin to gather and form small clumps and clusters (floc). This is the first step of coagulation; and for greatest efficiency, there must be rapid, intimate mixing of the raw water and the coagulant. After the period of rapid mixing, the water should be slowly stirred in order that the very small newly formed particles may attract and enmesh colloidal particles, holding them together to form larger particles or floc. This might be called the second stage of the process (flocculation).

In the discussion above the term turbidity is used frequently and might infer to some that mud and color turbidity are the only foreign materials removed by coagulation and settling. It should be strongly emphasized that a great part of the job of bacterial removal or purification of water is carried out in this phase of the treatment process. Bacteria are enmeshed or trapped in the coagulated particles and settled out, just as are the colloidal particles of turbidity. The bacteria have a negative charge the same as clay particles.

The amount of alum required to cause good floc to form varies with the water and to a great extent depends upon the amount of turbidity and color to be removed. Generally speaking, the more turbid the water, the greater the amount of alum necessary. Finely divided, suspended particles are more difficult to coagulate than the coarser particles and require a larger quantity of alum for a given turbidity. Clear waters are usually much harder to treat than muddy ones. Presently the best way to determine the proper, most economical amount of chemicals to add is to make the "jar test".

For any chemical reaction to take place, there must be two materials or substances to react. Therefore, equally as important as the chemical coagulant (alum) is the chemical composition of the water to which it is added. Mentioned previously is the fact that the alum reacts with the alkalinity in the water. The larger the amount or dose of alum used, the larger the amount of alkalinity that must be present for a complete chemical reaction. Some alkalinity is present in the water as natural alkalinity. Most times only alum is added to the water since there is enough alkalinity present for the reaction to be completed and good floc formed. When enough alkalinity is not present, an alkali must be added to supply the proper amount. The alkali commonly used is lime (hydrated lime). At a few plants, it may be desirable to use another alkali known as soda ash (sodium carbonate) or caustic soda (sodium hydroxide). Soda ash or caustic soda does not increase the hardness as does lime. Compared with lime, soda ash is much more expensive and is, therefore, used rather infrequently. As a rule of thumb, one ppm (part per million) of alum requires 0.4 to 0.5 ppm of alkalinity (natural and added).

There is at least one pH zone or range for every water at which the best floc is produced in the shortest time with the least amount of chemicals. Coagulation should be carried out in this pH range.

It is important that the coagulating process be closely controlled in order to get the clearest water for the least expense of chemicals. This control is maintained largely through the regulation of the pH of the water. Each time the raw water changes, "jar tests" should be made and the new chemical dosage rate and pH range found and maintained. The pH tests are made continually (one hour intervals) by the better operators to detect any change in the flocculated water.

There are several things that can occur to change the pH and thus cause the water to cease to floc or to cause the efficiency to be reduced. When the pH test is made on the water in the flocculator and it is noted that there has been a change, efforts should be begun to determine the cause. One of two things may have happened:

1. The amount of chemicals being fed has changed. (Even the best feeders may vary or clog and thus change amounts fed.)

2. The raw water has changed. (Plants taking raw water directly from a stream are subject to sudden changes in raw water. This is less likely when water is obtained from a lake.)

There are a number of things which can cause the water of a stream to change. Foremost of these causes perhaps would be rains upstream. Waste products from industrial plants upstream discharged intermittently will cause changes in the chemical quality of the raw water. If the coagulating chemicals are not properly balanced there may be some floc but the water may have a smoky appearance from finely divided coagulant in suspension. This may go through the filters and afterwards settle out in the reservoir or mains. Also, under such conditions, if water is passed through the filters before post-lime for corrective treatment is added, a floc may form in the clearwell. This is known as after-floc.

In the settling basin, a small or medium sized floc that appears dense and drops or falls out rapidly is desired. Observation of the water both as it leaves the basin and on the filter will reveal if much of the floc remains unsettled, thereby putting an unnecessary burden on the filter. Observation. along with the pH and jar test, is required to produce the best water most economically. Detailed explanation of the jar test is given at the end of this Appendix. Even casual observation will indicate that a very fine floc does not settle well. Likewise, a large, fluffy floc will continue to hang near the surface and travel the entire length of the settling basin.

A true experience related here as an illustration of "tests" versus "eye" operation, may prove informative and interesting:

At a certain plant in Alabama the operator desired a salary increase which the Board denied but made a counter offer as money was a little hard to come-by. The Board indicated that if operating cost could be reduced then funds thus saved would be available for salary raises. The operator decided that chemical cost could probably be reduced. He requested the State Health Department engineers to visit his plant and work with him. The engineers went, carrying along a piece of rather unwieldly test equipment. This equipment was a laboratory stirrer used in making the "jar tests". The engineers and operator discussed his operating procedures and the engineers observed a good floc, good settling and a fairly clear water on the filters. A good bacteriological and chemical quality water was being produced. The chemical cost was running approximately $625 per month. Jar tests were made and they soon indicated there were at least two chemical doses which produced a good settleable floc and clear water. If a very large chemical dose was fed, the settled water was clear; but the jar tests also indicated that there was a very low dose which produced a water as good. The high dose was approximately equivalent to that which the operator had been feeding at the plant. The chemical feed machines were calibrated and set for the low dose, as indicated by the jar test, and a good water was produced on plant scale. The chemical cost at the plant the following months averaged about $125. Did the operator get his raise?

The above story is true and numerous others could be related. It illustrates what might be possible at many plants and small savings should be possible over a period of time at any plant, if it is operated on a scientific practical basis.

ALUMINUM SULFATE (ALUM) $Al_2(SO_4)_3$

Aluminum sulfate was one of the earliest coagulants used in water treatment and is now the most extensively used. It is made by reacting bauxite ore with sulfuric acid. The normal strength of liquid alum is adjusted to 8.3% while the strength of dry alum is 17%.

Dry alum is shipped in bulk truck loads or bags while liquid alum is shipped in tank trucks of 3600 to 5000 gallon capacity.

One point to make in the feeding of alum, wet or dry, is that it can be fed either by weight or volume. The type of feeder would be determined by quantities to be fed, accuracy of feed, storage capacity, feeder location, etc.

Liquid alum must have special storage tanks made of fiberglass resin, stainless steel (very expensive), or steel tanks fitted with a rubber or vinyl lining. Feeding is normally done by a chemical feed pump which takes suction from the alum storage tank. Another method is by gravity flow through a flow meter at a controlled rate. The materials used in handling, feeding, and storage of liquid alum would most likely consist of PVC (polyvinyl chloride), glass fiber reinforced plastic, lined steel, rubber, teflon, or neoprene.

Some interesting characteristics of alum which should be noted are:

1. The "pH" of liquid alum should be around 2.5
2. Freezing point at 4°F
3. Liquid alum 8.3% Al_2O_3, Dry alum 17%
4. Chemical formula: $Al_2(SO_4)_3 \cdot 14H_2O$
5. Molecular weight 599 (Al_2O_3 = 102)
6. Appearance is clear to a white to a brown creamy color depending upon where the clay or bauxite came from, whether the filtration process was performed in the manufacture of it.
7. Crystalization occurs due to evaporation at high Baumes (60°F) (concentration too high) and atmospheric temperature
8. Weight of one gallon liquid alum weighs between 10.8 lbs. to 11.1, which would equal approximately 4.9 lbs. to 5.3 lbs. of dry alum

One other term that must not be overlooked, for it is a must in ordering liquid alum and feeding, is the word, Baume. A loose definition may be stated that Baume is a measurement in degrees to determine the specific gravity of a liquid, and the specific gravity is the weight relation of a liquid to that of water. All Baume readings are taken or based on 60°F; therefore, adjustments would need to be made when taking a reading. For this we would refer to a temperature conversion table usually supplied by all manufacturers. From this table we can convert the reading back to 60°F, finding the correct specific gravity, giving the pounds of available alum, and now can adjust feeding rates. Also, as pointed out above, crystallization in storage tanks, feeding equipment and solution lines may occur due to a high Baume. This usually occurs at a Baume of 36.6. If this does happen, adjustment to a lower Baume will place the crystals back into solution, making feeding a possibility. This lowering is called dilution and is done simply by the addition of water to lower the Baume.

Laboratory tests involving liquid alum would consist of reading the temperature in °F, the Baume, the specific gravity, residual aluminum, and dosage requirements by means of a standard jar test procedure, the last being the most important. However, in making stock solutions used in the jar test, you still need the other tests, excluding residual aluminum, to make up an alum solution where 1 ml would equal 1 mg of alum per 1000 ml raw water. Here is a method of determining the amount of liquid alum used in making up a stock solution.

First read the Baume and temperature, and then refer to conversion table to find specific gravity. Also, this should give you the percentage of Al_2O_3, and the weight per gallon. From this you proceed as follows:

Example

Baume 35.9, sp. gr. 1.3290, % Al_2O_3 8.23
1 gram dry alum per 1000 ml = 171 mg/l Al_2O_3
171 mg/l = molecular weight Al_2O_3(102) ÷ molecular weight $Al_2(SO_4)_3$(600)
1 ml liquid alum = 1000 ml x sp. gr. x percentage
1 ml liquid alum = 1000 x 1.329 x .0823 = 109.39 mg Al_2O_3
ml liquid alum = 171 ÷ 109.39 = 1.6 ml
1.6 ml liquid alum + 1000 ml distilled H_2O = stock solution
Stock solution = 1 ml + 1000 ml raw water = 1 mg or 1 ppm Al_2O_3

With regard to safety and the use of liquid alum, it should be noted that liquid alum is not dangerous to handle, but is an acid salt and must be handled with care. It can be irritating to the eyes and mucus membranes. If a spill occurs, the liquid alum becomes very slick due to rapid evaporation and should be washed off at once.

Many new surface water treatment plants are being designed to use liquid alum. Some existing treatment plants have converted their facilities to handle and feed liquid alum. This is economical only when the plant is large and thus uses a large amount of alum and when the plant and equipment alteration cost is not too great. Below are some advantages of using liquid alum but it should be pointed out that there is no mention of one type alum treating water better than the other type. The advantages are mainly the result of buying a liquid in large quantities and thus not having to handle a dry chemical.

1. Liquid alum costs less (approximately 1.1 cent per pound).
2. Less labor is required to unload, store and convey.
3. Dissolving operations are eliminated.
4. Less storage space required.
5. Greater accuracy in measurement and control with liquid alum is possible.
6. Nuisance and unpleasantness of handling dry alum is eliminated.
7. Housekeeping is simplified.

CHEMICAL FEEDERS

In the discussion on coagulation, it has been stressed that the proper amount of coagulant must be added to the incoming raw water. It is important that the proper amount be added so that the water is properly processed and to reduce cost. In addition to adding the coagulant, other chemicals may be added for taste and odor control (carbon), corrosion control (lime or soda ash) and removal of iron and manganese ($KMnO_4$). The theory or discussion on these phases of water treatment will be given later, but the feeders will be discussed here.

Feeders may be divided into two general types, (1) dry feeders and (2) solution feeders. The Alabama State Health Department's requirements with regard to dry chemical feeders are:

"Coagulating chemical feeding devices may be either dry feed or solution type. If dry feed machines are used, they should be installed in duplicate or repair parts stocked for speedy repair. They should have a range of 2 to 50 ppm for alkali and 4 to 100 ppm for alum and other coagulants."

DRY CHEMICAL FEEDERS

The dry feed type machine adds the chemicals either by direct weighing or by volumetric additions. In general all the dry feeders are equipped with a hopper or bin to hold the chemical. Various types of mechanisms are utilized to proportion the chemical from the hopper to a small stream of water, which conveys it to the incoming raw water. The proportioning mechanism can be adjusted so as to vary the amount of chemical fed. Most adjustments to regulate the operation of the dry feed equipment are made manually; however, machines are being made which are controlled automatically; that is, the dose varies in proportion to the quantity of water being treated.

Some chemicals feed much easier than others due to their own peculiar physical and chemical qualities. One important factor is the hygroscopic properties of certain chemicals which cause stickiness or dampness, etc. The "arching" properties of certain chemicals (lime) interferes with constant feed through a hopper. This is overcome by using vibrators, agitators, etc., in the hopper. One very important item aiding in correct and continuous feeding is proper storage of the chemicals (in a dry, cool place) and to have them as fresh as possible. Do not secure too large a supply at one time. There are various qualities of chemicals and an appropriate quality should be obtained. For example, it is not a good idea to secure ordinary building lime from a local hardware. There are better qualities available. In addition to this, the hardware or building supply house may have had the lime stored for long periods of time so that it has become lumpy and partially hard.

The design and construction of the conduit for conveying the chemical solution to the point of application should receive thought in order to save maintenance trouble later. Locate the feeder as near as practical to the point of application. Use troughs rather than pipe or hose when possible. Hose is preferable to pipe; however, if either is used it should be a relatively large size, (normally not less than 1¼ inches). The material should not be subject to attack by the chemical. Hose should be installed so that it does not sag. Facilities should be provided to flush these lines. They should be flushed daily with water under pressure to prevent clogging.

Separate lines, conduits, or troughs must be used for each different chemical. Two chemical solutions cannot be conveyed in the same conduit successfully. The chemical solutions should not discharge at the same point. Any chemical such as lime, soda ash, chlorine or fluoride should be added before the alum is added.

When a new plant is being designed or an old one modified, attention should be given to both the location of the feeders and the chemical storage room. The feeders should be located in a separate room from other operating equipment, but opening into or near the chemical storage room in order to reduce labor. Convenient access should be provided so that incoming chemicals can be unloaded and stored with a minimum of effort. The freight elevator should be so located that the chemicals can be handled with ease.

EXACT FEEDING REQUIRED

The accurate and uniform feeding of chemicals is probably the most important feature of plant operation and the one that requires most attention from the operator. He should exercise almost constant supervision over this feed to see that no clogging of orifices, sticking of float valves, or other things, vary the uniform flow of the chemical. It is absolutely essential that all parts of the feeding devices be absolutely clean at all times, as an accumulation of spilled chemicals, deposited sediment, or corrosion deposit will impair the accuracy of the feed. The best index of his operation to a person inspecting the plant is the condition of his/her chemical feeding devices.

CALIBRATION OF DRY CHEMICAL FEEDERS

After it has been determined by coagulation tests how much chemical, either alum or lime, is required to properly coagulate a water, the chemical feeder must be adjusted to deliver the chemical at the selected rate. If an attempt is made to adjust the feeder by guess, in all likelihood too little or too much of the chemical will be added. Thus a chart or curve should be prepared which will show how much chemical, expressed in weight and parts per million, the machine will feed at each setting. On the majority of the feeders the range of the scale for settings is usually 1 to 10. The determinations of how much chemical is fed at each setting is known as calibration.

The feeding mechanism of each type of dry chemical feeder is adjustable. Most types can be adjusted in two ways, one by changing the opening (throat) through which chemicals are fed, and the other by changing the length of stroke. Although a scale is provided on the stroke adjustment mechanism, usually in numbers as 1, 1½, 2, etc., all new feeders must be calibrated to determine the amounts of chemicals that will be delivered at each setting. It must be remembered that a calibration for one throat opening will not hold true for another opening. Consequently, if the throat opening is changed the machine must be recalibrated.

The value of calibration lies in the fact that the chemical dosage may be easily changed to the desired amount by referring to the calibration chart or curve, changing the length of stroke accordingly.

Calibration of a dry chemical feeder is a fairly simple procedure requiring a set of scales and a watch. The scales should be capable of weighing very small amounts and should be graduated in grams. The watch preferably should have a sweep second hand. The procedure itself consists of collecting the chemical discharged by the machine at a particular setting or stroke length in a given time. In general, the weight of chemical discharged by a short stroke is very small and the collection time should be for two or three minutes and the average amount per minute calculated. When the stroke is lengthened, the time can be shortened to one minute and, with extremely long strokes or high setting to one-half minute. For simplicity of calculation, the delivery of the feeder should be converted to weight per minute. From this, the parts per million of chemical fed to the water at a given setting can be determined more easily.

The calculations involved in calibrating a dry chemical feeder are illustrated by the following examples.

(1) Assume a plant has a capacity of 400 gallons per minute which is equivalent to 576,000 gallons per day.

At a stroke length setting of 3 and medium throat opening, assume that the chemical feeder delivers 22 grams per minute. Grams must be converted to pounds by dividing the number of grams by 453.6 (453.6 grams equal one pound). Therefore, the rate of delivery is 22 divided by 453.6 = 0.048 pounds per minute.

The following formula is used for determining parts per million when the weight of chemical used and the number of gallons of water treated in a particular time are known.

$$\frac{\text{Number of pounds of chemical used} \times 120{,}000}{\text{Number of gallons of water treated}}$$

$$= \text{parts per million (ppm)}$$

Substituting the value 0.048 pounds of chemical added per minute and 400 gallons of water treated per minute in this formula, a result of 14.4 ppm is obtained.

$$\frac{0.048 \times 120{,}000}{400} = 14.4 \text{ ppm}$$

Another way of arriving at ppm is:

$$\frac{\text{Grams per minute} \times 265}{\text{Gallons per minute}} = \text{ppm}$$

(This is possible because the formula:

$$\frac{\frac{\text{grams/minute}}{453.6} \times 120{,}000}{\text{gallons per minute}} = \text{ppm}$$

is simplified by dividing 120,000 by 453.6 and obtaining the constant, 265).

As mentioned on the previous page, the weight of chemicals delivered in one minute at various settings of the stroke length should be determined and converted to parts per million. This information should be recorded in such a manner that it is easily available to the operator. One of the best ways of doing this is to prepare a chart and calibration curve which can be framed and hung near the feeder. Such a chart or a calibration curve is shown in Figure 2. It should be remembered that it is necessary to check the calibration of any feeder at intervals, at least every six months. Certainly when there is a change in chemicals, such as changing from ground alum to rice alum, the feeder should be recalibrated.

SOLUTION FEEDERS

Some of the newer and larger plants in the state use liquid alum or caustic soda. These solutions and others such as hydrofluo silicic acid (fluoride) and potassium permanganate are fed by means of a solution feeder. In most cases the solution feeder will take suction directly from the chemical storage tank. The operating parts of the feeder/are usually made of a strong plastic or stainless steel to resist any corrosive effect of the solution.

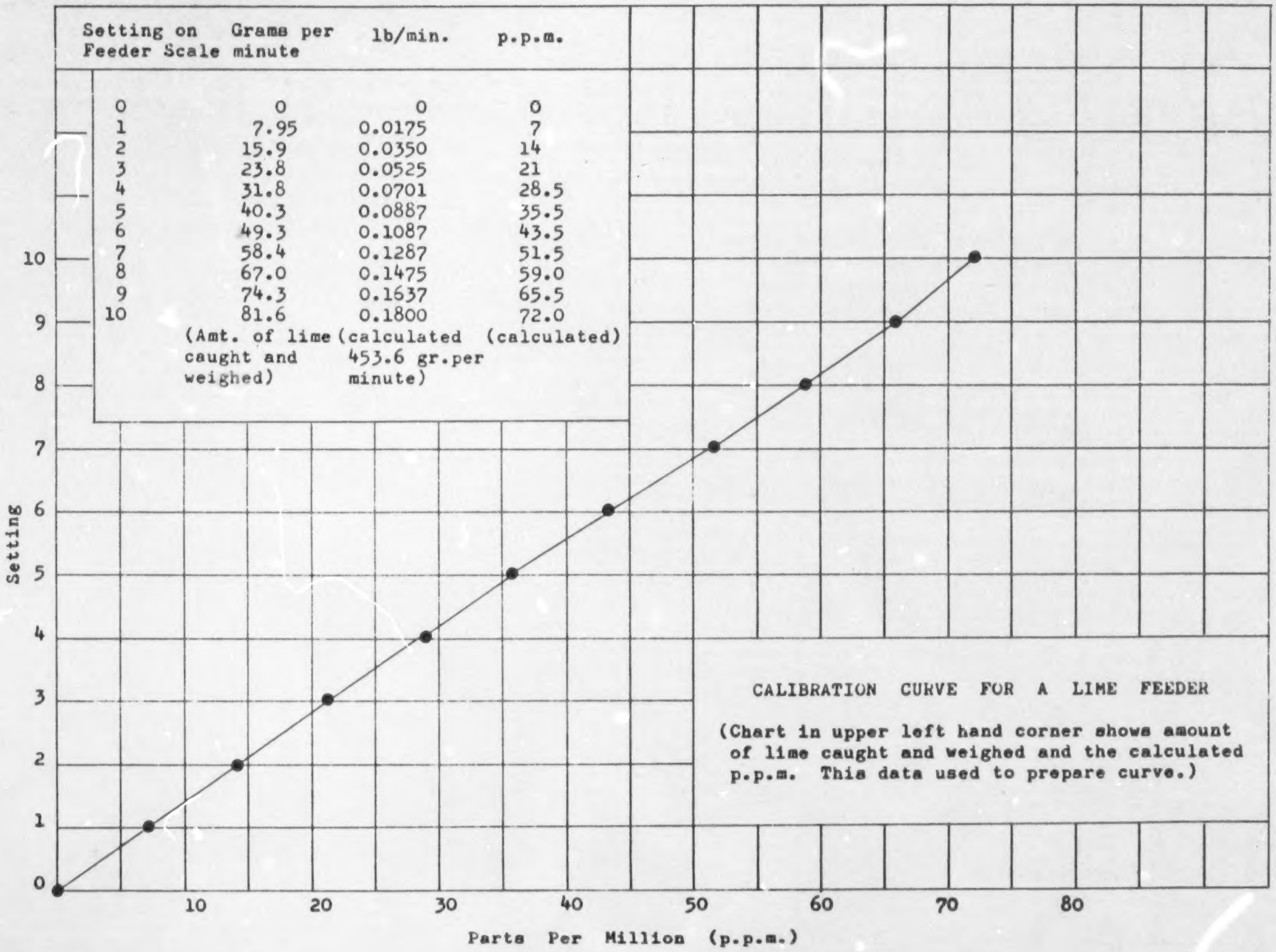
The most common type of solution feeder is the diaphram pump. This pump is designed to provide a solution at pressures up to 125 psi. A range of capacities is obtained by changing the pulley ratio between an externally mounted frictional motor and the input shaft of the pump drive mechanism. A closer adjustment of the pump delivery may be obtained by varying the stroke length by means of a screw adjustment during operation.

Another type solution feeder available is the positive displacement pump. Due to high cost, this type feeder is not frequently used. No matter what type solution feeder is to be used, the operator should study the operations instruction book which comes with the equipment and be thoroughly familiar with it.

The capacity range of a solution feeder is specified for the size of feeder being used but the machine should be calibrated to enable the operator to be able to make the needed adjustments quickly. This is best accomplished by measuring the volume discharged at the chemical application point. This can sometimes be difficult, and since there is usually no large discharge pressure involved, the chemical discharge can be measured at the feeder. A calibration curve for each pulley setting should be constructed. Information for this curve is found by measuring the discharge per time unit at different stroke length settings. Converting this discharge to a feed rate and plotting this versus the stroke setting gives the operator a curve he can pick a feeder setting from at the feed rate needed.

The calculations involved in calibrating a liquid chemical feeder are illustrated by the following examples.

Figure 2



Setting on Feeder Scale	Grams per minute	lb/min.	p.p.m.
0	0	0	0
1	7.95	0.0175	7
2	15.9	0.0350	14
3	23.8	0.0525	21
4	31.8	0.0701	28.5
5	40.3	0.0887	35.5
6	49.3	0.1087	43.5
7	58.4	0.1287	51.5
8	67.0	0.1475	59.0
9	74.3	0.1637	65.5
10	81.6	0.1800	72.0
	(Amt. of lime caught and weighed)	(calculated 453.6 gr.per minute)	(calculated)

CALIBRATION CURVE FOR A LIME FEEDER

(Chart in upper left hand corner shows amount of lime caught and weighed and the calculated p.p.m. This data used to prepare curve.)

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Assume a plant has a capacity of 2 MGD which is approximately 1400 gpm. The chemical feeder is operating at a certain rate of strokes per minute.

1. Assume a stroke setting of the machine of 12 strokes per minute. At a particular stroke length, say 1-¼", 116 ml of liquid alum by measurement was delivered for one stroke. This means that 12 x 116 or 1393 ml/min. is being delivered. In terms of gallons,

   $$\frac{1393 \text{ ml/min}}{3785 \text{ ml/gal}} = .368 \text{ gal/min}$$

   This is .368 gal/min x 1440 min/day = 530 gal/day. Liquid alum is only 8.3% effective alum so only 530x(.083) = 44 gal/day of alum is fed at this rate. For a 2 MGD plant, this is

   $$\frac{44}{2,000,000} = 22 \text{ ppm of alum}$$

   By adjusting the length of the stroke the amount of liquid alum changes. Another method of accomplishing the same thing would be to convert the concentration of liquid alum to dry alum weight (about 5.3 pounds per gallon) and using the relationship discussed in dry feeders of:

   $$\frac{\text{number of pounds of chemicals used x 120,000}}{\text{number of gallons treated}} = \text{ppm}$$

2. Assume a similar pump at a lower setting (4 strokes per minute) is going to be used to feed hydrofluosilicic acid (H_2SiF_6) to provide a fluoride concentration of one ppm at the same plant. Two gallons of full strength fluoride are needed for a 1 ppm concentration

   $$\frac{2 \text{ gallons}}{2,000,000 \text{ gal}} = 1 \text{ ppm}$$

   The H_2SiF_6 is 80% strength so $\frac{2 \text{ gal}}{.80}$ = 2.5 gal. needed. But this full strength H_2SiF_6 is only 23% effective fluoride so $\frac{2.5 \text{ gal}}{0.23}$ = 10.9 gal/day is needed. This is $\frac{10.9}{1440}$ = .0755 gal/min. or (3785 ml/gal) (.0755 gal/min) = 286 ml/min. This is $\frac{286}{4 \text{ strokes/min}}$ = 71.5 ml/stroke needed. The machine could be adjusted to give this amount by trial and error.

A calibration chart should be drawn showing the feed concentration at the normal plant rate or capacity. This would enable adjustments to be made easily and rapidly.

COAGULANT AIDS

Some waters are difficult to coagulate by the addition of aluminum or ferric salts. Almost all coagulant aids are very expensive and care must be taken to add only the proper amount.

Activated silica is a negatively charged sol which, when added to water from the flash mixer, forms long-chain inorganic polymers and attracts the positively charged metallic hydroxide sol created by the addition of alum. This attraction improves the formation of floc. This material is normally used in connection with treatment of very clear water that is hard to flocculate.

Polyelectrolytes are high molecular weight, long-chain, organic polymers with a multitude of ionizable sites along the chain length. When this material is added, it tends to reduce the stability of colloidal systems. This occurs because the ionizable sites dissociate to form charges on the polymer and the positive charges attract the negatively charged colloid particles forming large floc which will readily settle.

One use of coagulant aids has been as an emergency treatment of water which might have received inadequate treatment in the settling basin and flocculator. This might be the result of a severe change in raw water quality or a malfunction of an alum feeder. In a properly operated plant, these conditions should not go unobserved for such a length of time that a coagulant aid is needed to assist the filter in treating the water. The use of coagulant aids just before filtration can cause clogging of the filter and thus short filter runs. High velocity backwash may be required to adequately clean the filter if this occurs and this high velocity might upset the filter media.

JAR OR COAGULATION TEST

As previously stated, the coagulation of the water is influenced by a number of factors and there is no way of calculating the chemical dose needed from the turbidity, color, alkalinity or pH value of the raw water. The plant can be operated by observing the turbidity and the floc being formed and adjusting the chemical feeders until a good floc is obtained. This may be a very expensive dose. The most practical way to find the proper dose is by making the jar test. At one plant a zeta meter is used to aid the jar test.

Originally, a jar test was made by using large fruit jars or gallon jugs. Varying doses of alum and lime were added and each jug shaken a number of times. This caused the test to be quite laborous and unpopular. In order to secure definite results an apparatus known as a laboratory stirrer is needed. Laboratory stirrers may be secured from Phipps & Bird, Inc., Richmond, Virginia; Omega Machine Company, Providence, R.I.; Taulman and Company, Atlanta, Georgia; or other manufacturers.

In addition to the stirrer, beakers (glass laboratory vessels) which hold one liter are needed. The chemicals (usually alum and lime) needed in the test are the same as is being used in the plant. It is not necessary to know the purity (strength) of the chemicals, as they are the same as will be used in the full scale operation of the plant. All results of the tests will be in parts per million or milligrams per liter. Should the equivalent dose in grains per gallon be desired this figure can be obtained by dividing parts per million by 17.1.

The first step in making the test is to prepare a stock solution of the alum and lime. Stock solutions should always be prepared of such a chemical strength that when one ml. (milliliter) of solution is added to the portion of sample being used it is equal to the addition of one ppm of the chemical. Distilled water should be used in preparing all stock solutions. Below are given examples showing how to prepare stock solutions to be used with different amounts of water samples. The first example is the most commonly used.

Example 1. Using one liter (1,000 cc. or ml.) raw water samples in the test.

One ml. of water weighs 1 gram. One liter contains 1,000 ml. and therefore its weight is 1,000 grams. If one gram of chemical is added to one liter of distilled water, the concentration of chemical is one part in 1,000 parts or 1,000 parts per million. One ml. of this solution would contain $\frac{1}{1000}$ gram or one milligram of the chemical. The addition of one ml. of the stock solution to one liter of the raw water to be tested results in a concentration of the chemical in the water samples of $\frac{1}{1000}$ gram in 1000 grams ($\frac{1}{1000} \times \frac{1}{1000}$) or one part per million.

Example 2. Using two liter raw water samples in the test.

In this case the amount of water sample is twice that used in Example 1; therefore, the weight of chemical to prepare one liter of stock solution will be two grams rather than one gram. One ml. of this stock solution added to two liters of raw water sample will be equivalent to adding one ppm of chemical.

Example 3. Using one gallon water samples in test.

One gallon contains 3.785 liters (approximately 3.8 liters). Thus the weight of chemical needed to prepare one liter of stock solution will be 3.785 grams.

Example 4. Preparation of one gallon of stock solution for use in treating one gallon water samples.

In this case the problem is to prepare one gallon of stock solution to use in treating one gallon raw water samples of water. One gallon contains 3.785 liters or 3785 ml. Since it is desirable to have the strength of the stock solution such that one ml. added to the sample portion is equivalent to one ppm, 3.785 x 3.785 or 14.33 grams of chemical must be used as there are 3785 ml. of distilled water and 3785 ml. of sample.

Lime is not completely soluble, hence a true solution cannot be made. It is necessary, therefore, to agitate the stock solution (or suspension) every time it is used in order to secure the desired dose.

PROCEDURE

Fill the beakers or jars to the mark measured for the amount of raw water sample to be treated (usually 1000 ml). Next run an alkalinity test on the raw water. If the raw water has an alkalinity greater than 15 ppm, alum alone is tried first. If prelime has to be fed, it is usually best to hold the amount of lime constant and vary the alum dosage. Next, with pipettes, measure various amounts of the stock solution into small beakers. Record in tabular form the dose of each stock solution added to each beaker & the way in which it settles. Many times after the first test is made and the approximate dosage is determined, the test is run again, varying the dose slightly from the best one observed in the first test to find out exactly how much chemical is required. The smallest amount of chemicals which gives the best settling floc should indicate the rate at which the incoming raw water should be treated.

Suggested Table

Jar Number	Alum ppm	Alkali ppm	pH Value	Results of Floc
1	10	0	6.9	No floc
2	15	0	6.8	Fair
3	20	0	6.7	Good
4	25	0	6.6	Good
5	30	0	6.4	Fair
6	35	0		No floc

Figure 3

"Suggested Table" (Figure 3) shows the data collected from a jar test and indicates how such a record may be prepared. When the stirring operation is begun, the lime (if needed) is added first then the alum is poured from the small beaker into the raw water sample. Operate the stirring equipment as nearly as possible to

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represent the conditions in the mixing and/or flocculating facilities at the plant. For example, where there are a rapid mixer and flocculators, the stirrer should be operated very fast for 45 seconds to one and one-half minutes, depending upon the retention time in the mixer, and slower for 6 to 10 minutes to simulate the flocculator.

COAGULATION WITH ZETA POTENTIAL

As more and more plants are designed for high rate filtration, and as the requirements for the physical and chemical quality of the water improve, additional methods must be used in order to properly treat the water. Two methods are in use in many plants in the United States today----(1) Zeta Potential, and (2) Pilot Filters. Zeta potential will be briefly discussed here.

As mentioned previously in this Appendix, colloidal particles, which help make up a large amount of turbidity in the water, have an electrical charge which is usually negative. As you know, like electrical charges will repel (as the like poles of magnets); therefore, as long as there is a sufficient charge on the colloidal particles, they will not come together to form floc and settle out.

The magnitude of the electrical charge surrounding the colloidal particles is called the Zeta potential. Zeta potential could also be defined as the repulsive force that keeps turbidity in suspension. An instrument that measures the magnitude of the electrical charge is called a zeta meter.

Since the charge on most of the colloidal particles is negative, then the coagulating chemical must be one that will put enough positive charges in the water to reduce the charge surrounding the colloidal particle.

For many years, it was thought that the basic coagulation chemical reaction indicated that the aluminum was hydrolyzed to the insoluble trihydrate ($Al(OH)_3$):

$$Al_2(SO_4)_3 + 3\ Ca(HCO_3)_2 \rightarrow 2Al(OH)_3 \downarrow + 3\ CaSO_4 + 6\ CO_2 \uparrow$$

However, this could not be the case because all the compounds formed in the above reaction are stable and would have no positive electrical charges to put in the water to neutralize the zeta potential of the turbidity colloids.

A recent scientific explanation is that one or more of the following multivalent soluble ions is formed:

(a) $Al_6(OH)_{15}{+++}$;
(b) $Al_8(OH)_{20}{++++}$;
(c) $Al_{13}(OH)_{34}{+++++}$.

These are soluble ions with a positive charge that are capable of reversing the negative charges of colloids. In other words, these are the important types of ions that coagulate turbidity rather than the formation and settling of the aluminum trihydrate ($Al(OH)_3$) compound.

After the above multivalent soluble ions are formed and the zeta potential surrounding the colloidal particles approaches the zero point, then the water is gently stirred (flocculated) in order to bring the many particles in close proximity of each other. When the electrically neutral particles come close enough together, then a molecular attraction force (called Vander Waal's Force) causes the particles of turbidity to come together. When enough of these particles join together into a large enough "floc" particle, then they settle out.

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CHEMICAL COAGULATION*

Large, heavy particles suspended in water will settle out quite readily, but smaller lighter particles settle very slowly, or in some cases not at all. Chemical coagulation consists of treating the water with certain chemicals to bring the non-settleable particles together into larger, heavier masses of solid material, which are then comparatively easy to remove. These masses are called "floc".

The practice of coagulation includes several operations, discussed below, because each of them involves different considerations. The first step in the process consists of *feeding* one or more chemicals to the water. Secondly, *mixing* is necessary, so that the chemicals are rapidly and evenly distributed, and all parts of the water will be equally well treated. During the mixing, or immediately afterward, certain chemical reactions occur. Some of the reaction products are insoluble, and they begin to precipitate as solid particles. The term *coagulation* is sometimes used to refer specifically to these chemical reactions and the beginning of the formation of the floc. Since the same word is also used to mean the total overall process up to sedimentation, it must be used rather carefully to avoid confusion. The final step in the process is *flocculation*, which refers to gentle agitation of the treated water for a period of time. This favors the collision of the small floc particles with each other and with the other suspended particles in the water. This causes them to stick together and grow into large, readily settleable masses. Again, the word flocculation may have two slightly different meanings. In theoretical discussions it may mean the growth of the floc particles. In plant practice it more often refers to the gentle agitation of the water which brings about this growth. *Sedimentation*, generally, follows flocculation, but is usually considered a separate process and not a part of chemical coagulation. It refers to the removal of the floc from the water by settling in basins especially designed for the purpose.

Detailed steps in the coagulation-sedimentation process have changed during the past several years, and for this reason confusion still exists regarding such terms as "mixing", "coagulation", "flocculation", "sedimentation" and others. Insofar as this discussion is concerned, "mixing" refers to the rapid, uniform distribution of a coagulant or other chemical in the water being treated before chemical reactions have occurred to any marked degree. "Coagulation" refers to the physical-chemical changes which take place between the soluble coagulant and the alkalinity in the water to form precipitated or incipient floc. "Flocculation" refers to the gentle agitation of a coagulant treated water for an appreciable period of time to foster the coagulation reactions and to provide conditions favorable for the flocculent material to meet and adhere together into large floc masses. The term "sedimentation" refers to the settling of floc in basins specifically designed for that purpose.

*Source: Manual of Instruction for Water Treatment Plant Operators, New York State Dept. of Health, Health Education Service, P. O. Box 7283, Albany, N. Y. 12224, Chapter 7.

Chemicals Used for Coagulation

Several different chemicals may be used in coagulation, depending upon the characteristics of the water being treated. In some waters a combination of two or more chemicals produces better results than any one chemical does alone. Usually, to decide which chemical or combination should be used, it is necessary to perform coagulation tests in the laboratory. These tests are described later in this Appendix.

Aluminum sulfate. Filter alum or merely alum is the most commonly used coagulant in water treatment. It may be purchased in solid form, or as a solution called "liquid alum". "Activated alum" contains about nine percent sodium silicate. This improves the coagulation in some waters. "Black alum" is alum containing activated carbon.

Sodium aluminate. "Sodium aluminate" is a much more alkaline compound than alum. When added alkalinity is found to be necessary for the best coagulation, sodium aluminate may be substituted for all or part of the needed alum dose.

Iron salts. The iron salts most commonly used as coagulants include ferric sulfate, ferric chloride, ferrous sulfate (also known as copperas), and chlorinated copperas. These compounds often produce good coagulation when conditions are too acid for best results with alum. Sometimes the particles are best removed under acid conditions and iron compounds give better results.

Coagulant aids. These materials do not produce coagulation when they alone are added to the water, but improve the results when they are used with a coagulant. Several of the chemicals used in this way are either acid or alkaline compounds and are used to adjust the pH for best coagulation. These include lime, sodium carbonate (known as soda ash), sodium hydroxide (known as caustic soda or lye), and sulfuric acid.

Other coagulant aids have little or no effect upon the pH, but enter into the coagulation process. Some of them are believed to act by providing solid surfaces where the first formation of floc can begin more easily. Materials used for this purpose include certain clays (such as kaolinite and bentonite), sodium silicate, and activated silica. These materials may also add weight to the floc and improve sedimentation.

Still others are believed to react with the fine particles in the water and assist in bringing them together into larger and heavier masses for settling. Coagulant aids of this type have molecules which consist of long chains of atoms, with many reactive groups scattered along the length of the chain. These are mostly artifical products. Those known as "synthetic polymers" or "synthetic polyelectrolyes" are entirely man-made. Others are manufactured from starch or cellulose. The molecules of starch and cellulose are already long chains and, in manufacturing, the reactive groups are added.

Chemical Application

There are a number of chemical compounds which are used as water coagulants, but aluminum sulfate, commonly called alum, is most generally used. Alum is granular powder-like material which is readily soluble in water and is easily applied as a solution or as a dry material. Reactions between alum and the natural constituents in various waters are influenced by many factors and it is good operation practice to determine the required amount of coagulant to be added by trial or experiment.

Normally, the best coagulation results are obtained with an alum dosage from 10 to 50 ppm, with the optimum dosage being obtained through the use of the so-called "jar test". This technique consists of adding known amounts of a coagulant into several jars of the water to be treated, gently stirring this mixture for a definite period of time, and observing the quality and settling characteristics of the floc.

Sometimes a coagulant dosage is given in grains per gallon (gpg) or more conveniently, in pounds per million gallons (lbs/mg). The relationship between these units of measure is illustrated in Table I which may be used to convert dosages from one system of units to another.

TABLE I

Converting from	*Multiplication factor to get*		
	gpg	*ppm*	*lbs/mg*
Grains per gallon (gpg).	1.00	17.1	143
Parts per million (ppm).	0.0585	1.00	8.34
Pounds per million gallons (lbs/mg). .	0.007	0.12	1.00

Examples:

1. Express a dose of alum of 0.6 gpg in terms of ppm and lbs/mg.
   0.6 x 17.1 = 10.26 ppm. 0.6 x 143 = 85.8 lbs/mg
2. Express a dose of 30 ppm in terms of gpg and lbs/mg.
   30 x 0.0585 = 1.75 gpg. 30 x 8.34 = 250.2 lbs/mg.
3. How many pounds per day of alum would be required to treat 2.5 mgd with a dose of 25 ppm?
   25 x 8.34 x 2.5 = 521.3 pounds per day.

Coagulants are usually fed by a machine which can be set to deliver a known amount during a certain period of time. The machine may be calibrated in convenient units, such as pounds per twenty-four hours. There are available many types of machines or chemical feeders. A brief description of some types will suffice. Some machines feed alum in the dry form either directly or indirectly to the water treated. Usually the dry alum is fed at a known constant rate to a small tank, or "solution pot", where it is suspended in water. The resulting slurry is added to the water. Solution feeders, on the other hand, consist of tanks to hold solutions of known strength and a means for applying known amounts of the solution to the water being treated.

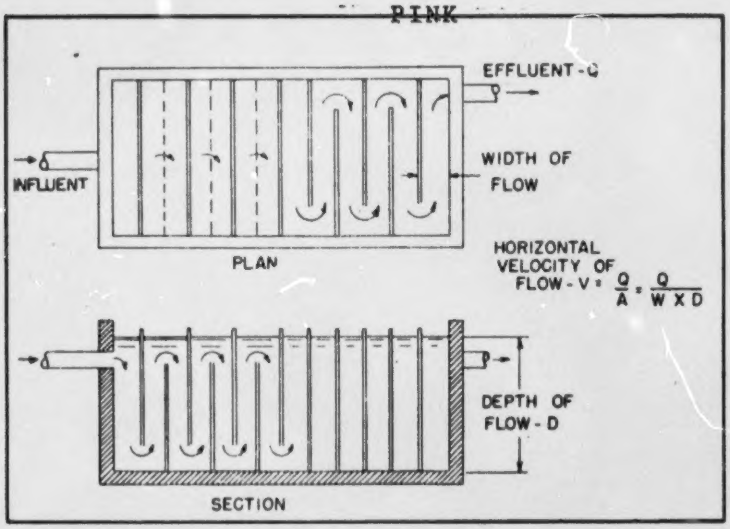
"Mixing" is necessary to provide rapid and uniform dispersion of the alum throughout the water being treated. It may be done either mechanically or hydraulicly in special tanks, in sections of other tanks, or in piping systems. The fundamental principle is the violent agitation of the water with the applied chemical for a short period of time. This may be accomplished either by motor driven propellers or baffled channels to create turbulent flow conditions. Contrary to popular belief, centrifugal pumps in themselves do not provide good mixing as their efficient performance depends on a minimum amount of turbulence as the water passes through them. Valves, elbows, and other devices normally used in conjunction with pumping operations may cause adequate turbulence for satisfactory mixing.

Flocculation takes place following mixing, usually in a single tank with a detention period between fifteen and forty-five minutes. During this period the water is gently agitated to make floc particles. Bacteria and suspended solids meet and adhere together into large masses. The physical-chemical mechanism by which this is accomplished is very complex and beyond the scope of this discussion. Suffice it to say that only by this gentle agitation will the coagulated particles become large enough to settle readily. Generally, soft waters of low mineral content coagulate best over a narrow range of pH values, between pH 5.8 to 6.4, whereas harder waters, such as those from the Great Lakes, coagulate quite readily at pH values of 6.8 to 7.8. Under normal conditions the reaction between alum and the alkaline materials usually found in a water supply are quite effective for the removal of turbidity and this is accompanied by the absorption of a moderate amount of color. Water with high concentrations of color, greater than about 30 units, requires that the coagulation reaction be in the "acid zone" of pH value about 5.0 to 6.0. In this range a more complex action takes place to form what is termed a "color floc" instead the conventional aluminum hydroxide floc.

Normally, there is sufficient alkalinity present in the water to complete the coagulation process. At times, it is necessary to add alkalinity, in the form of lime or soda ash. When it is necessary to coagulate in the "acid zone" in order to remove color, an acid may be added to obtain the proper pH value. Such waters are quite corrosive and must be subsequently treated with an alkali to prevent corrosion from occurring in the distribution system.

The gentle agitation that is essential in flocculation may be accomplished either hydraulically or mechanically. The most common hydraulic method is the well known baffled basin, in which the water flows "around the end" or "over and under" baffles. Unfortunately, these baffles are not readily adjustable, and the degree of agitation depends upon the rate of velocity of the water. Since the optimum degree of agitation varies, this lack of flexibility is undesirable. Flow velocities between baffles from 0.3 to 1.0 foot per second usually give reasonably satisfactory results.

Figure 1

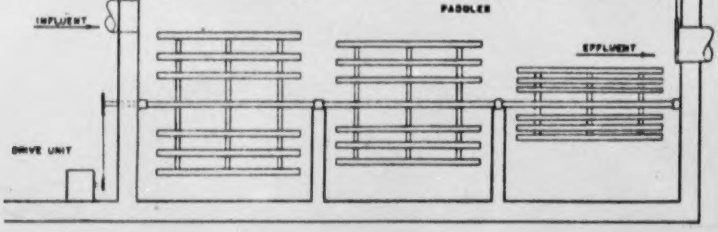

EFFLUENT-Q
INFLUENT
WIDTH OF FLOW
PLAN
HORIZONTAL VELOCITY OF FLOW - $V = \frac{Q}{A} = \frac{Q}{W \times D}$
DEPTH OF FLOW - D
SECTION

BAFFLE FLOCCULATION BASIN

The present trend is away from the use of baffles in flocculation tanks. In their place, mechanically driven equipment is installed to provide the required agitation. The most common type of equipment consists of power-driven paddles or agitators so proportioned that appreciable but not undue agitation is produced. The speed of rotation can be controlled and the equipment operated to give the optimum degree of agitation regardless of the characteristics or quantity of water being treated or type of coagulant being used. The paddles may rotate on either a vertical or horizontal axis, and either longitudinally or at right angles to the direction of flow. Sometimes, several paddles are operated in series and the first paddle is rotated at a greater speed than the succeeding ones. Modern practice suggests sudden violent agitation as the coagulant is added with a gradual decrease in turbulence in each successive step through the entire process of "mixing", "flocculation" and sedimentation". Figure 2 illustrates a mechanically driven paddle type flocculation basin.

Essentially, the control of flocculation basins consists of regulating the amount of coagulant that is added as well as the degree of agitation in order to secure the most effective floc formation with a minimum coagulant dosage. At times, it is desirable to use coagulant aids in order to produce a floc which settles readily. It is good practice to equip the outlet end of the basins with submerged lights in order to permit observation of the floc that is formed. Properly coagulated water should show visible floc in clear water, like snowflakes in clear air, as contrasted to cloudy water or foggy air.

Figure 2



MECHANICAL FLOCCULATION BASIN

CORROSION AND SCALE CONTROL*

(CORROSION)

Importance

The control of corrosion is important not only to the continuous efficient operation of water plants and systems but also to the delivery of properly conditioned water for use in other operations such as heating and air conditioning. Wherever corrosion is minimized, there is an appreciable reduction in maintenance and operation costs. Corrosion is often a complex engineering problem because of the many combinations of causes, influencing factors, and effects that are possible in any situation. While the principles and methods of corrosion control discussed herein pertain specifically to water works facilities, they are in general applicable to all utility systems.

a. *Economics.* Corrosion control is a very important factor in reducing the long-range cost of maintenance and operation of real property facilities. A survey conducted in 1950 by the National Association of Corrosion Engineers placed corrosion losses in the United States at six billion dollars a year. The Air Force share of this huge loss runs into millions of dollars, and emphasizes the urgency of an effective control program. While such a program cannot eliminate all corrosion, it can reduce it by a high percentage.

b. *Operation and Safety.* Corrosion, especially when it attacks installation utility systems, can interfere seriously with the accomplishment of the base mission by causing plant shutdowns for repair and rehabilitation, and by wasting such critical materials as copper and nickel. Corrosion causes leaks in mains and domestic piping, resulting in waste of water. Products of corrosion cause clogged pipelines and equipment and a reduction in pipeline flow capacity. Corrosion also causes water to become turbid and imparts a color which will stain fixtures and clothing in laundering. All of these are evidence that corrosion is taking place. Corrosion is also a threat to safe operation and may cause toxic hazards to installation personnel.

Theory of Corrosion

a. *Definition.* Corrosion is now commonly accepted to be an electrochemical reaction in which metal deteriorates or is destroyed in contact with its environment or surroundings, such as air, water, or soil. Corrosion is essentially the same in all metals. Such refined metals as steel are relatively unstable and tend to change back to the stable compounds and oxides of their natural ores.

*Source: Maintenance and Operation of Water Plants and Systems, Department of the Air Force, Air Force Manual #85-13, June 1968, pp. 209-228, 235-36.

b. *Electrochemical Reaction:*

(1) The term electrochemical reaction means that chemical changes and an exchange of electrical energy take place at the same time. Thus, wherever corrosion occurs, there is a flow of electric current from the corroding portion of metal into the electrolyte or conductor of electricity, such as water or soil. The point at which current flows from the metal into the electrolyte is called the anode, and the point to which current flows from the electrolyte is called the cathode.

(2) The electric current flowing from the metal carries particles of metal with it. The charged particles are called ions. As the metallic ions are dissolved in the electrolyte, they are exchanged for hydrogen ions, which flow to the cathode and tend to deposit a hydrogen film on the cathode. The metallic particles carried from the parent metal are usually changed to oxides as they move into the electrolyte and are deposited as corrosion products, most often near or on the anode. In the case of steel, these oxides form the familiar pattern of rust.

Corrosion and Scale Relationship

Corrosion and scale are associated problems, but their effects and cause should not be confused. The essential effect of corrosion is to destroy metal; scale tends to clog open sections and line surfaces with deposits. The products of corrosion often contribute to scale formation and aggravate the problem of its treatment. Thin scale may provide a protective coating against corrosive attack on metal; on the other hand, thick scale (as in hot-water heating boilers) may accelerate corrosion at other points in the equipment by decreasing heat transfer and causing overheating. Both corrosion and scale formation can and frequently do occur together as a result of a number of outside factors. For instance, the presence of slime affects both corrosion and scale formation. In general, it increases scale because the microorganisms use up the materials used for scale treatment. Slime also forms a sticky surface that makes it easier for deposits to adhere to metal. This same coating on metal may increase corrosion by preventing the penetration of treatment agents.

Types of Corrosion

a. *Localized.* Localized corrosion is confined to a small area of a metallic system and is evidence of corrosion caused by an electrochemical reaction. It is characterized by pitting or the formation of cavities that develop under tubercles (as noted in cast iron water mains). Localized corrosion results from such causes as small pits in metal, differences in the composition or finish of metal, or differences in stress. Local cell action occurs when areas on a metal surface in an electrolyte (conductor of electricity such as water) have different potentials. Pitting is usually more serious than uniform corrosion in that it indicates that corrosion is heavily concentrated in one area, thus permitting rapid corrosion penetration and failure.

b. *Uniform.* Uniform corrosion is an attack over a whole area of a metal surface, exemplified by the rusting of sheet iron and the tarnishing of silver.

c. *Compositional.* Compositional corrosion causes a change in the structure of materials. An example of this type of corrosion is dezincification, in which a zinc alloy (especially yellow brass) loses zinc that is replaced by a spongy residue of some less active constituent such as copper. Other examples are graphitization (in which the metallic iron constituent of gray cast iron is changed to corrosion products, leaving graphite) and hydrogen embrittlement (in which hydrogen entering the metal causes it to lose ductility or its ability to be deformed).

d. *Corrosion Fatigue.* This type of corrosion occurs when a metal is subject to repeated stresses. Well rods, rotating shafts and similar objects frequently fail from corrosion fatigue. Prior to failure, the surface indications of corrosion are often very minute.

Factors Influencing Corrosion

a. *Biological.* Chief among biological factors influencing corrosion are the activities of bacterial organisms.

(1) *Soil Corrosion.* Biological corrosion results from chemicals produced by soil bacteria. Iron structures buried in soil are attacked by sulfate-reducing bacteria, which thrive in wet, poorly drained soil with low aeration. These bacteria are anaerobic (thrive in the absence of oxygen) and require adequate amounts of sulfate in the soil. They reduce sulfate to sulfide by utilizing hydrogen, thus producing hydrogen sulfide which reacts with iron. The principal product of such corrosion is black iron sulfide. Iron bacteria, on the other hand, are aerobic (require oxygen). They do not appear to attack iron directly but, by consuming ferrous iron in solution and eliminating it as hydroxide, form deposits that may cause tubercles on the metal surface. Iron bacteria in some cases are also responsible for red water.

(2) *Fouling.* Fouling is caused by organisms, chiefly in sea water, that form slimy deposits and underwater growths on metal. Important among these marine organisms are algae, sponges, mollusks, and barnacles. Bacterial slimes result in an uneven deposit on the metal and under certain conditions prevent normal protective film formation.

b. *Chemical.* Chemical factors influencing corrosion include:

(1) *Mineral Salts.* Mineral salts dissolved in the water will settle out under certain conditions when the concentration is above saturation values. The scale thus formed will insulate the metal

and protect it from corrosion. High concentrations of soluble sulfates or chlorides increase the corrosiveness of water and require higher chemical dosages for corrosion protection.

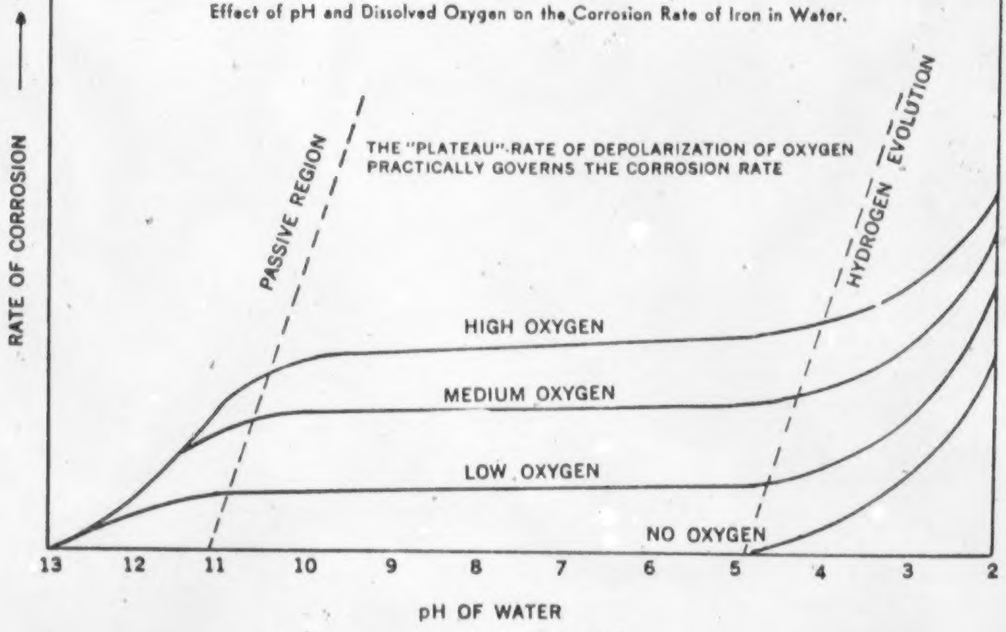
(2) *Dissolved Oxygen.* Dissolved oxygen is a depolarizing agent; that is, it removes accumulated hydrogen from metal surfaces submerged in water, allowing oxygen to combine with the metal and form oxides. This results in cavities or pitting. For example, air accumulating in the top of a vertical hot water tank causes corrosion on the water side. This corrosive attack is increased by the presence of carbon dioxide.

(3) *Carbon Dioxide.* When dissolved in water, carbon dioxide forms a weak acid, known as carbonic acid. This acid combines chemically with iron to cause corrosion. The products of the reaction are ferrous becarbonate and hydrogen gas. This form of corrosion is frequently encountered in hot water lines and steam condensate return lines and associated equipment. Failures occur most frequently at threaded joints.

(4) *Hydrogen Ion Concentration (pH).* Hydrogen ion concentration (symbol pH) is a measure of the acidity or alkalinity of a liquid. Corrosion is generally more rapid in liquids with a low pH (therefore acid) than in alkaline solutions. Hydrogen ions tend to form hydrogen gas on metal surfaces and excape from corrosive water solutions, removing the protective gas film from the surfaces, as acidity increases. As indicated in Figure 1, the corrosion rate is lowest at a pH of 14, increases to a pH of 10, and remains essentially uniform until a pH of 4.3 is reached. When considering the effect of pH alone, the rate and type of corrosion in water containing dissolved oxygen falls into 3 pH zones:

Figure 1

Effect of pH and Dissolved Oxygen on the Corrosion Rate of Iron in Water.



(a) *Acid.* When pH is less than 4.5, corrosion is general over the entire iron surface and hydrogen gas is formed on a relatively small portion of the metal surface. Rate of corrosion increases rapidly as the pH decreases. Rust or tubercles do not form on the metal.

(b) *Neutral.* Rust always forms when pH is between 4.5 and 9.0, the range of most natural waters. Rate of corrosion is nearly dependent on pH, but increases with temperature and concentration of oxygen. Corrosion occurs mainly in fairly well-defined areas which become more definite as pH increases. Other parts of the iron surface are attacked very little; at low oxygen concentration, hydrogen is liberated at these parts. Even in the absence of oxygen, iron corrodes in water at a slow rate to form black iron oxide, magnetite, Fe_3O_4, and hydrogen.

(c) *Alkaline.* When pH is 9.0 or more, corrosion decreases rapidly with increase in pH. Rust coating is denser and more adherent. Severe pitting occurs under the tubercles.

(5) *Vapor.* Vapors from certain acids, such as sulfurous, hydrochloric, and nitric acid have a corrosive effect on metals, particularly in the presence of oxygen and moisture. Hydrogen sulfide attack on such metal surfaces as window frames in sewage treatment plants is an example of this action.

c. *Physical and Mechanical.* Physical and mechanical factors influencing corrosion include:

(1) *Erosion.* Erosion is the wearing away of metal by mechanical action, as a pump impeller may be eroded by sand carried in water. This abrasive action is not a form of corrosion but may accelerate the destruction of metal because corrosion products are more easily eroded than the metal itself. Liquids and gases, especially at high velocities, can cause this mechanical wear and wash away protective coatings on metal surfaces.

(2) *Cavitation.* Cavitation (sometimes called cavitation erosion) is the damage caused by the formation and collapse of gas bubbles (generally air) of cavities in liquid where the metal and liquid are in contact. For example, as liquid flows over a turbine blade, a suction or low pressure area forms and creates bubbles in the liquid. When the bubbles break, they cause pitting in the metal at that point.

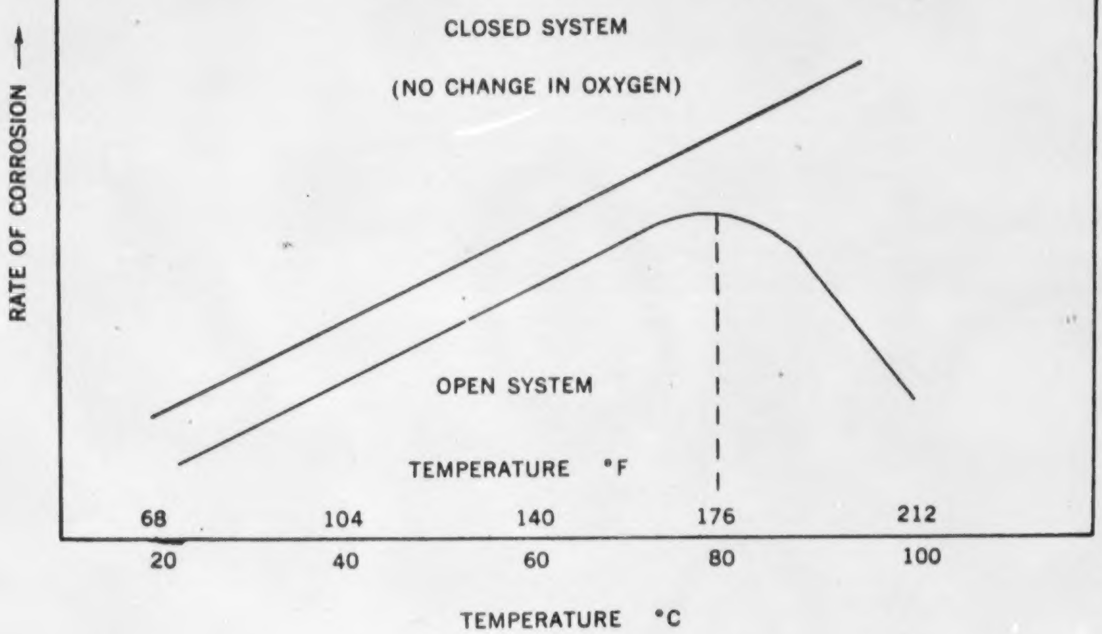
(3) *Metal Porosity.* Porous areas in metal may be caused, for example, in casting. Sandholes or blowholes left in the metal not only weaken its structure but allow deeper penetration by the seepage of corrosive liquids.

(4) *Pressure.* Pressure alone does not appear to affect the rate of corrosion; however, it increases the solubility of oxygen in water and adds stress to metal in pressure vessels, both of which are conditions favorable to corrosion.

(5) *Temperature.* Changes in temperature increase or decrease the rate of corrosion reaction, the effect depending on other conditions. For example, in a closed system where no variation in dissolved oxygen can occur, corrosion is accelerated by increases in temperature, as shown in Figure 2. In an open system, the maximum rate of corrosion is at 80 degrees C. Above that temperature, oxygen is expelled from water because of its low solubility at high temperatures, and consequently the rate of corrosion drops sharply. Normally, the higher the temperature, the more rapid the chemical reaction and the higher the corrosion rate.

Figure 2

Effect of Temperature on Corrosion of Iron.



(6) *Time.* The length of time metal is exposed to a corrosive environment may affect the rate of corrosion. The initial rate is generally high, but may decrease as natural protective rust coatings are formed on the surface of the metal. (Corrosion rate is generally measured in thousandths of an inch per year.)

(7) *Velocity.* As liquid velocity increases, the liquid tends to wash away protective coatings and supply more oxygen or carbon dioxide to the area of contact, thus stimulating corrosion. Motion of liquid may have a uniform effect over the whole metal surface and prevent localized corrosion, or it may set up turbulence at certain points to where it will cause pitting. For iron, at constant pH and

dissolved oxygen values, the rate of corrosion reaches a maximum at 1.3 feet per second (fps); it then declines sharply up to 3 fps, and increases beyond that velocity because of combined erosion and corrosion effects. High localized velocities may cause corrosion by impingement attack.

d. *Electrochemical.* When a difference in potential exists between two parts of a metal structure which is also in or in contact with an electrolyte, a flow of electric current occurs. The potential may be set up by two dissimilar metals, by a difference in temperature, amount of oxygen, or concentration of the electrolyte at two points of contact with the metal. The anode is the point where current flow is from the metal to the electrolyte; it is here that corrosion occurs. The cathode, which is usually not attacked, is the point of current flow from the electrolyte to the metal. Some forms of electrochemical corrosion are:

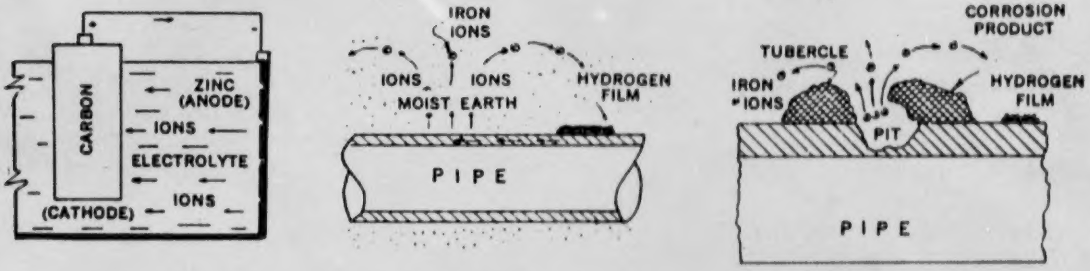
(1) *Concentration Cell.* A concentration cell is an electrolytic cell producing a current because of a difference in the concentration of the electrolyte or oxygen at the anode and cathode. Crevice or contact corrosion is typical of this reaction. For example, if pipe is laid in soil that is not uniformly compacted or is backfilled improperly, crevices in the backfill may cause one area of the pipe to be aerated more than another. This creates a difference in potential which sets up a corrosive action.

(2) *Galvanic Action.* When two dissimilar metals are brought into contact or are connected by a conductor, an electric current is created. For example, in a flashlight battery, carbon and zinc are placed in an electrolyte and connected externally with a wire. In producing current, zinc (the anode) goes into an invisible solution in the form of zinc ions. After continued use, most of the zinc disintegrates until it breaks into pit holes. This wastage of zinc is referred to as galvanic corrosion (Figure 3). This behavior is entirely analogous to "galvanic cells" found along a pipeline with dissimilar metal couplings. For instance, a brass valve in an iron pipe will act as a galvanic cell, with the brass being the cathode, the current passing through the soil (the electrolyte in this case) to the brass valve, returning through the metal to the anode area (Figure 4a). Where current leaves the pipe, it will carry metal with it, and thus the pipe will be corroded. Another example is where a piece of new pipe is connected in an old line. The new line will always be the anode and corrode away (Figure 4b). When two dissimilar metals are located in a moist soil (electrolyte), one acts as anode and the other as cathode according to their relative position in the Electromotive Series. Table 1 shows the relative position of metals of common usage. If any two metals in this series are placed in an electrolyte and externally connected, the metal listed nearest the top of the series will be anodic and will corrode. Galvanic action can also occur between identical metals in different electrolytes. For instance, dissimilarity between soils in which a pipeline is laid can result in a flow of electricity. The degree of galvanic corrosion is determined largely by the type of soil in which the line is buried. Corrosion is generally

most rapid in acid soils. Soils are generally classified by resistivity, as high resistance or low resistance soils. A pipeline passing through soils of different resistivities will corrode in the low resistance soils.

Figure 3

Principles of Galvanic Corrosion.


CARBON
ZINC (ANODE)
IONS
ELECTROLYTE
IONS
(CATHODE)
IRON IONS
IONS
IONS
MOIST EARTH
HYDROGEN FILM
PIPE
CORROSION PRODUCT
TUBERCLE
IRON IONS
HYDROGEN FILM
PIT
PIPE

(a) BATTERY ACTION **(b) IONS IN MOTION START CORROSION** **(c) CORROSION PROCEEDS WITH PITTING & TUBERCULATION**

Figure 4

Corrosion Caused by Dissimilar Metals.

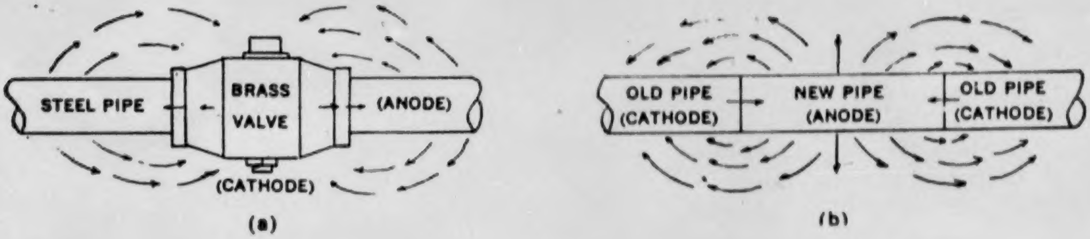

STEEL PIPE
BRASS VALVE
(ANODE)
(CATHODE)
(a)
OLD PIPE (CATHODE)
NEW PIPE (ANODE)
OLD PIPE (CATHODE)
(b)

TABLE 1. ELECTROMOTIVE SERIES FOR COMMON METALS

Anodic End
Magnesium
Zinc
Aluminum
Carbon Steel
Copper Steel
Cast Iron
Lead
Tin
Nickel
Brass
Copper
Bronze
Copper Nickel
Monel
Silver
Gold
Cathodic End

(3) *Electrolysis*. Corrosion caused by stray direct-current (d-c) voltage is an example of electroylsis, or the production of a chemical change in an electrolyte as the result of the passage of electricity. These stray currents usually develop from electric railways where the current, instead of returning to the generator through the rails, leaves the rails and follows the path of least resistance to nearby underground metal structures such as pipelines and cables. Corrosion occurs where the current leaves the pipeline to return to the generator. Because of the reversal in direction of current flow, alternating current (a-c) does not cause this action.

e. *Metallurgical*. Metallurgical considerations affecting corrosion include stresses set up in metals through the fabrication process or by such operations as welding, soldering, and riveting. Such stresses may make metals more subject to corrosion, and the combination can result in stress corrosion cracking.

Corrosive Environments

Of the corrosive environments, water (including steam and vapor) is the chief cause of corrosion and, in its many forms, the most common electrolyte. Soil and atmosphere are the other important corrosive environments.

a. *Water:*

(1) Water is rarely found in a pure state; it may carry a variety of substances, such as minerals, salts, acids, organic matter, oil, and grease. Such impurities either cause or accelerate corrosion. On the other hand, impure water may gradually deposit mineral or organic matter that builds up a protective coating to retard corrosion of metal, whereas chemically pure water will not and therefore tends to be more corrosive.

(2) In electrochemical corrosion, water is the most frequent conductor of electric charges. Salt water has an especially high rate of conductivity and is very corrosive. Direct factors in creating a corrosive water environment may be the introduction of hydrogen sulfide by sewage discharges, the addition of brines to cooling waters, demineralizing water, an increase in acidity from industrial wastes and acid mine waters, and similar local conditions.

b. *Soil*. Soil conditions are responsible for the exterior corrosion of metal structures under or in contact with the ground. When two identical electrodes (such as two sections of a pipeline) are connected in dissimilar electrolytes (such as different soils), current flows from one electrode, through the electrolytes, to the other electrode, setting up a corrosion cell. Among the many causes of differential electrolyte cells in the soil are: non-uniform texture of the soil in contact with a tank or pipe; differential aeration or difference in the availability of oxygen at two points on the metal surface; two or more different types of soil, such as sand and clay, in contact with a metal object; presence of foreign matter, such as cinders, in the soil, and differences in the moisture content of the soil around a metal object. In general, salt-bearing soils, highly alkaline soils, clay, brackish marshes, and poorly drained soils are conducive to corrosion.

c. *Atmosphere.* In atmospheric corrosion, moisture is the main controlling variable. The action of salt air, hydrogen sulfide, sulfur dioxide, and other substances is much more rapid in moist than dry air. The presence or absence of certain foreign elements in the metal sometimes has a marked effect. Metals, such as zinc, aluminum, nickel, and copper, form a more or less protective coating which retards further corrosion.

General Methods of Corrosion Control

In a given situation, corrosion may have one cause or many; hence control measures must be planned to meet the specific conditions. Since it would be impossible to give specific measures covering all local problems, the following overall control measures are listed in sufficient detail to guide personnel in establishing corrosion control measures and solving specific problems.

a. *Designing against Corrosion.* Two steps are necessary in general design: to incorporate features that minimize corrosion at the outset; and to establish conditions that will make it easier to combat corrosion showing up in subsequent operations. Good design entails the selection of proper materials, the allowance of extra thickness of metal where it is known corrosion may occur, the proper protection of metals by protective coatings and linings, the installation of cathodic protection, and the provision of adequate water treatment devices. Good design provides for easy access to structures and equipment (such as handholes and ease of disassembly) for inspection, replacement, or remedial treatment and for suitable protective measures that may become necessary in the future.

b. *Selection of Materials:*

(1) *Proper Choice of Materials.* A basic step in the prevention of corrosion is the selection of the most resistant material for a given environment. (The term "corrosion resistant" may be misleading if applied to a material generally; corrosion resistance must always be measured in terms of a particular application in a particular corrosive medium.) Manufacturers usually offer a choice of materials for critical parts of equipment, such as pumps, that may be exposed to corrosion. Among light metals, aluminum, magnesium, and cadmium are relatively corrosion resistant, and among heavy metals, cast iron and lead. The nonferrous metals, such as copper and bronze, are generally more resistant to corrosion than the ferrous metals. Nonmetallic materials that may be substituted for metal in some applications are wood, concrete, and asbestos cement. Concrete, wrought iron, or asbestos-cement pipe are excellent for salt water and soft and corrosive waters but not acids. For corrosive water conditions cement-lined cast iron pipe should be used in lieu of plain cast iron. In good soils it is safe to use cast iron

pipe. In highly corrosive soils use of either cement-asbestos pipe, cement-lined and coated steel pipe, or coal tar enameled pipe is preferable. Copper piping for use as service connections and interior plumbing is highly resistant to corrosion. It should not be used, however, for very soft waters with large amounts of carbon dioxide. Brass piping containing 85 percent or more of copper is also very highly resistant to corrosive waters. Lead pipe is satisfactory for hard waters but should never be used with soft waters because of the toxicity of lead salts.

(2) *Avoiding Dissimilar Metals.* By using the same metal as far as possible throughout the piping system, galvanic action can be minimized. For instance, the use of brass valves or bronze fittings with iron piping should be avoided. (See Figure 4a.) Where it is necessary to use dissimilar metals use metals close together in the electromotive series. Coatings such as cement mortar or asphalt where the metals are pointed afford protection and minimize galvanic corrosion. Use insulating joints, sleeves, and bushings between dissimilar metals to prevent galvanic current flow. These are made from non-conductive materials such as neoprene, bakelite and nylon. Galvanic corrosion can also be reduced by inserting short sections of nonconductive piping, such as plastic or hard rubber piping, between the dissimilar metals.

c. *Cathodic Protection.* Whenever current leaves a metal surface located in an electrolyte, electrically charged particles of the metal go into solution, causing wastage of metal or corrosion. If the flow of this current can be stopped or reversed, theoretically the metal will not corrode. In cathodic protection, the current flow is retarded by passing an auxiliary current through the electrolyte to the metal.

(1) *Definition.* Cathodic protection is defined as the application of sufficient direct current to a metal object to prevent currents from leaving the anodic areas of the structure. The entire metal object then becomes negative or cathodic to its surroundings. For each environment, a certain minimum current must be applied to mitigate corrosion. The source of current is immaterial as long as the amount of current is adequate.

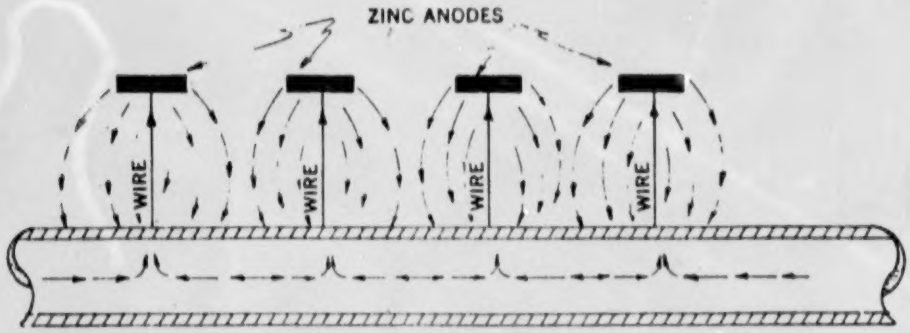
(2) *Methods.* The flow of protective current may be motivated by sacrificial anodes (galvanic method) and by impressed current.

(a) To achieve cathodic protection by the galvanic method, a sacrificial anode (a metal which is more anodic and that will corrode more readily than the object to be protected) is put in the electrolyte and electrically connected to the object. Electric current will then flow from the more anodic or corroding metal through the electrolyte to the metal for which protection is desired. The electrical connection between the two objects completes the circuit, allowing the current to return to the corroding metal. The corroding metal becomes the anode of a purposely established dissimilar metal cell, and the metal to be protected becomes the cathode.

Generally, the relative position of the metals in the electromotive series (Table 1) will determine the direction of current flow. The current from the corroding sacrificial anode is strong enough to prevent all current from leaving what were the anodes of the metal to be protected. Since the cathodic protection currents flow from the electrodes, they act as sacrificial anodes and disintegrate. The protected metal then becomes cathodic or negative to its surroundings. Metals such as magnesium and zinc which are higher in the electromotive series have been found effective sacrificial anodes when coupled with iron (Figure 5). Iron, on the other hand, is more anodic than copper and may be used to protect copper sheathings and structures. For large structures or piping systems requiring large currents, cathodic protection by the galvanic method is not economical and a direct current must be used.

Figure 5

Principle of Cathodic Protection.



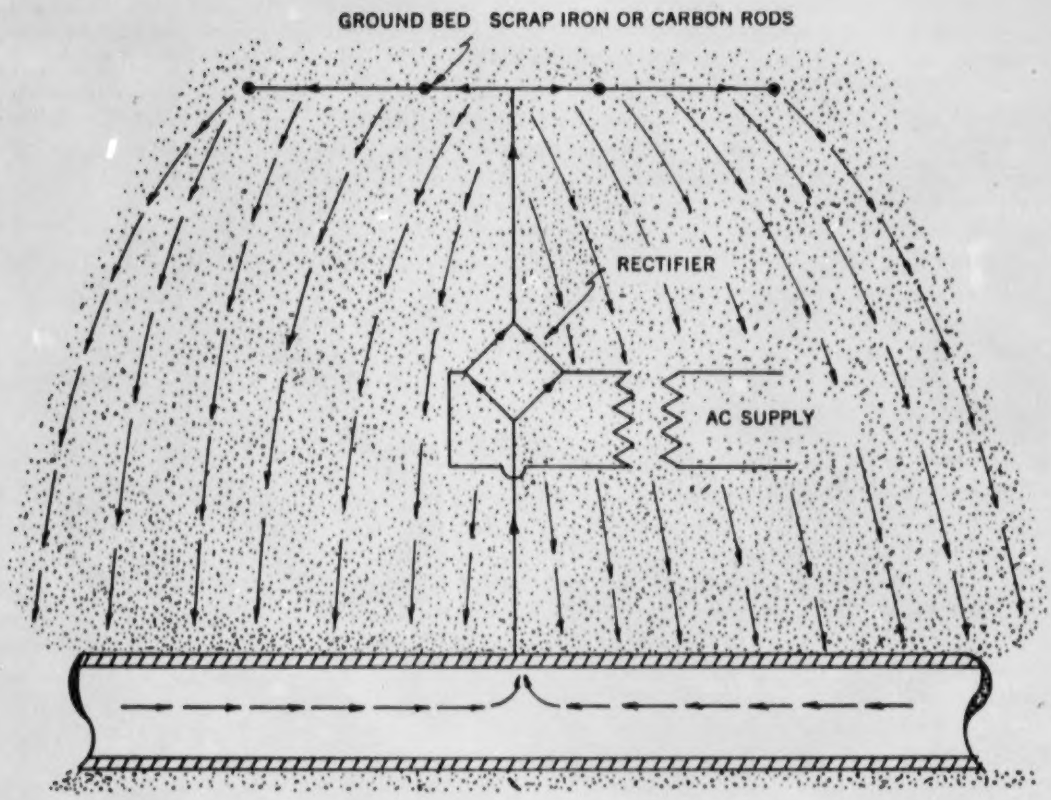
PIPE LINE OR OTHER BURIED METAL STRUCTURE

(b) To achieve cathodic protection by the impressed current method, a direct current from an external source such as storage batteries, wet or dry rectifiers, or motor gasoline engine generators must be used. The positive terminal of the generator is connected to iron, carbon, or graphite electrodes buried near the structure to be protected and the negative terminal of the generator source is connected directly to the structure. The external power source is used to force the current through the ground to the structure and then back to the power source through the negative terminal. This current neutralizes the normal galvanic current flowing from the structure into the ground. Figure 6 illustrates this form of cathodic protection.

(3) *Limitations*. Cathodic protection is limited in application to only certain structures and cannot be used to retard corrosion in general. For instance, the outside surface of a buried pipe

Figure 6

Cathodic Protection of Buried Pipeline Using External Power Source.



can be protected cathodically, but not the inside surface. In steel elevated water storage tanks, it is the inside surfaces of the riser and the submerged bowl that can be protected; protection of the outside of the tank from atmospheric corrosion must be by some other means. The application of cathodic protection for a few metal structures is discussed briefly below.

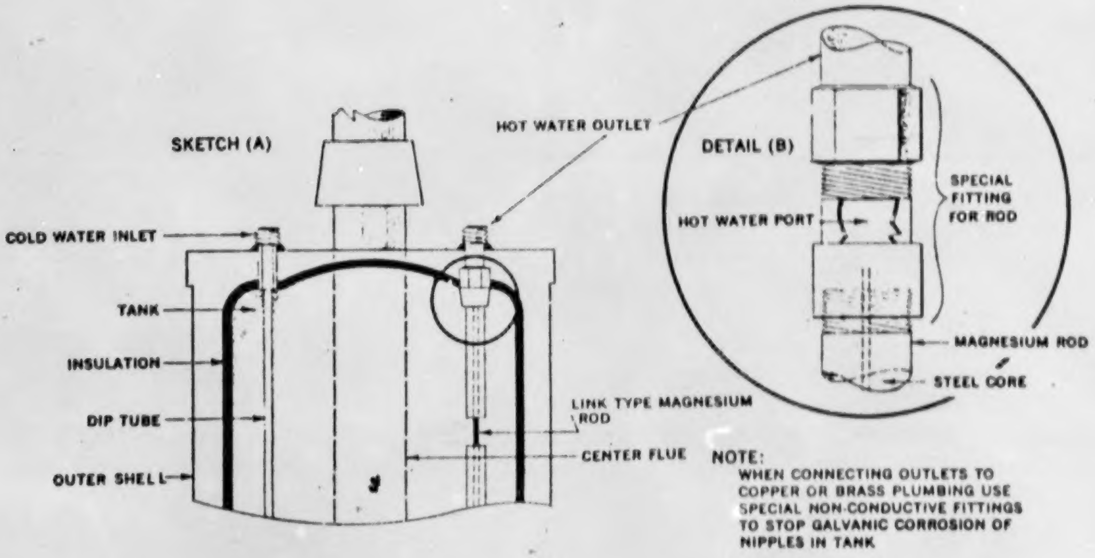
(a) *Buried Pipelines*. Buried pipelines are most commonly protected by cathodic means using an external power source. It is necessary, therefore, to force the current into the soil, and through the soil to the pipe line to be protected, and return it to the source. The principle difficulty is forcing the current into the soil. Contact is made with the earth by burying scrap metal such as rails or pipe, or carbon rods. Figure 6 shows a conventional installation. The current requirement to provide complete cathodic protection varies over an extremely wide range, but in general depends on the electrical resistance of the pipe coating, area of exposed metal, soil resistance, metal-to-soil potentials of cathodic and anodic areas, distribution of cathodic protection current and kind of metal in the structure. The major factor is the amount of current required in the area of bare or exposed metal. For example, on a mile of bare 12-inch pipe line with about 18,000 square feet of metal surface, 100 to 150 amperes of current might be necessary for adequate protection. In a mile of a well coated 12-inch line, a fraction of an ampere may be sufficient, which indicates the importance of a good coating. No general rule can be given, although for preliminary estimating purposes 0.003 amperes per square foot of bare surface is used for iron structures. After installation of cathodic protection, sufficient external current is applied until the pipe-to-soil potential is lowered to at least minus 0.8 volts as measured by a copper sulphate reference electrode, using a high resistance voltmeter. When this potential is maintained, corrosion and leaks on pipe lines will be greatly reduced or stopped.

(b) *Steel Elevated Water Storage Tanks*. Cathodic protection stops corrosion by flow of direct current to the tank wall which is reverse to the current created by corrosion. The equipment consists of a rectifier to produce a direct current from the alternating supply, control rheostats, an ammeter, and electrodes suspended from the tank roof into the water. The elctrodes, made of platinum or aluminum, are connected to the positive wire of a d-c circuit. The negative wire is connected to the tank bottom, completing the circuit for current flow from the electrodes to the tank surface. The induced electrical current flows continuously from the electrode to the tank. This current must be equal from all parts of the tank. Current density of about 5.0 milliamperes per square foot is necessary for bare metal; less is required if the tank is partially protected by paint. For even distribution of current density, it is essential to locate the electrodes at certain predetermined configurations, otherwise full protection for the entire submerged surface will not be secured. In wide diameter water tanks, it is necessary to locate a group or circle of electrodes near the tank side walls, with another group around the center area to protect the tank floor. For protection of the riser, a long electrode is usually suspended within a few feet of ground level. The current requirement to provide complete cathodic protection varies with a number of factors, the most important of which are the resistivity of water and the number, diameter, and spacing of the electrodes. For water with high resistance (low dissolved solids content), higher current densities are required than for water with low resistance. For example, midwestern waters with a specific resistance of 1200 ohms as compared to eastern seaboard waters with a specific resistance 10 to 20 times as great will require 1/20th of the power consumption for a comparable eastern tank with the same electrode arrangement. A great deal of emphasis is usually placed upon the corrosion occurring on the intermittently submerged surface of a tank, since as pointed out before only submerged surfaces can be protected cathodically. This and the problem of corrosion due to moist air on the surfaces above the high water line are usually dealt with by a coat of paint which extends to the mean water line when cathodic protection is applied. To reduce current requirements, the entire interior of the tank is usually painted.

(c) *Hot Water Tanks*. Hot water tanks are most commonly protected with metals which develop galvanic cells with the shell of the tank and therefore do not require an external power source. Magnesium rod anodes are extensively used. For small domestic hot water tanks, these anode rods are easily installed through the hot water outlet by using a special fitting. The rods come in link sections for tank locations where the headroom is limited and solid rods are provided for locations where there is at least four feet of headroom above the tank top. The rods are equipped with several types of fittings (3/4" nipples, plugs, elbows for electric heaters, etc.). These special fittings together with the magnesium rods can be furnished by commercial firms. Figure 7 illustrates use of special fittings. The anodes should be full length and come within a few inches of the tank bottom but not touch it. A good contact should be made at the hot water outlet to obtain continuous current flow from the rod through the water to control corrosion by cathodic protection. Tanks for corrosive waters need more than one rod or single rods of larger diameters or made of special alloy. The size and number of rods needed can be determined by certain calculations based on the total dissolved solids of the water and the temperature and diameter of the tank. Total dissolved solids content of a

Figure 7

Hot Water Tank Showing Magnesium Rods and Special Fittings.



NOTE:
WHEN CONNECTING OUTLETS TO COPPER OR BRASS PLUMBING USE SPECIAL NON-CONDUCTIVE FITTINGS TO STOP GALVANIC CORROSION OF NIPPLES IN TANK

base water supply is obtainable from the chemical analyses made annually by the U.S. Geological Survey, copies of which are available at each base. The following table presents a fairly good estimate for all practical purposes.

Diameter of tank in inches	Total dissolved solids of water in ppm	*Number of 0.84 (27/32") diameter magnesium anodes needed
12	100 to 500	1
12	50	2
18	200 to 500	1
18	100	2
18	50	3
24	200 to 500	1
24	100	2
24	50	4

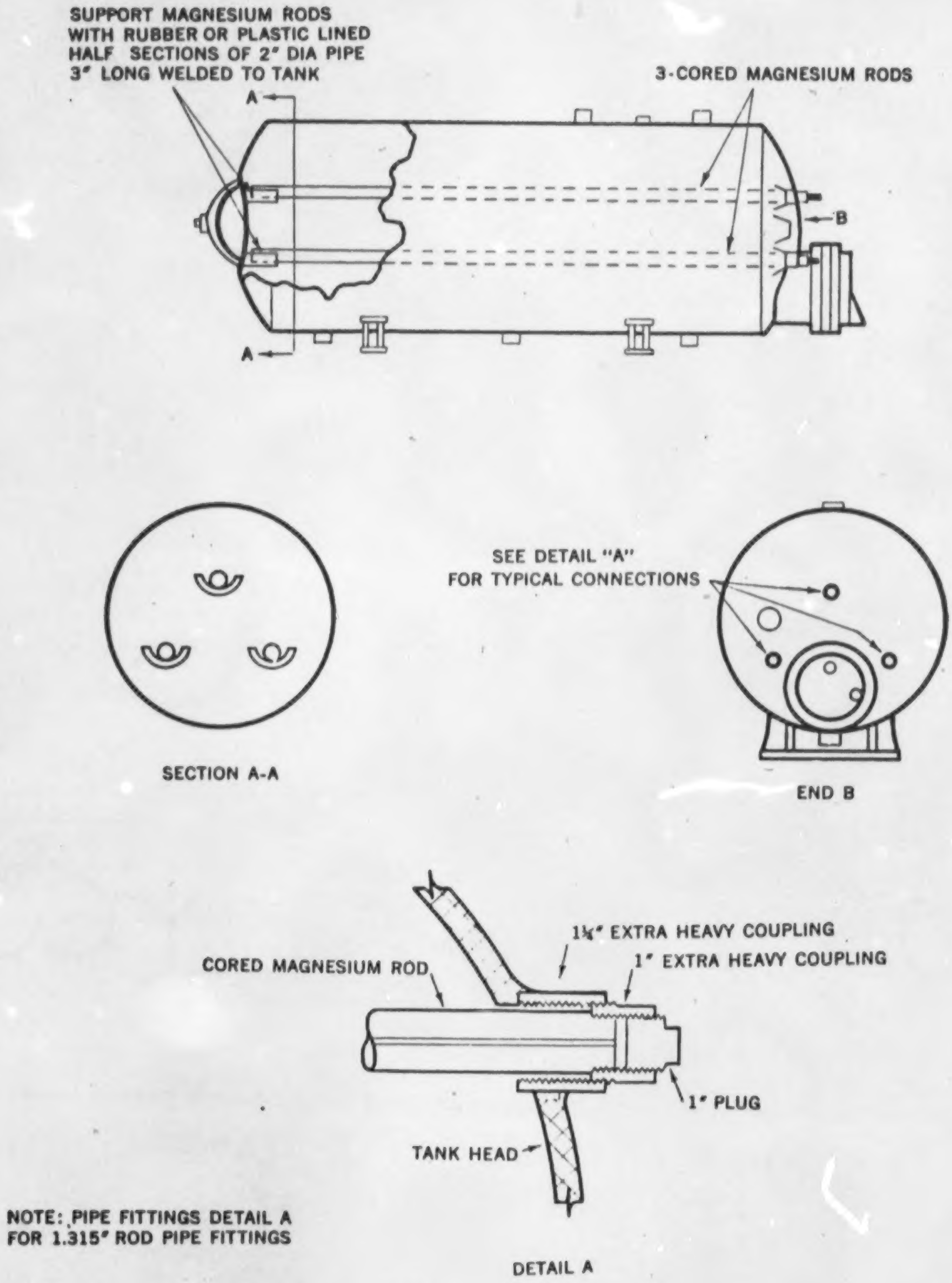
**Note: Number of anodes can be reduced by using larger diameter of special alloy rods. The anodes should be checked every year or two as the actual protective life of the anode is slightly less than the theoretical life. Once the anode rod is spongy in appearance and very thin it should be renewed.*

Diameter of Tank	Total dissolved solids of water in PPM				
	500	300	200	100	50
	Approximate life of magnesium rod in years				
	1.5	2.6	3.8	7.8	13
	Number of 1.315" diameter magnesium rods required				
30"	1	1	2	3	6
36"	1	1	2	3	6
42"	1	2	2	4	—
48"	1	2	2	4	—
60"	1	2	3	6	—
72"	2	2	4	—	—

Anodes should not be placed in tanks which are corroded to a point of mechanical failure. When new tanks are being purchased, consideration should be given to obtaining 10-year-warranty water heaters with glass linings or extra heavy galvanized steel tanks with built-in magnesium anodes. For larger hot water tanks (generators) the magnesium rods can be secured to the inside of the tank by welding ordinary pipe fittings to the tanks as shown in Figure 8. The size and number of magnesium rods needed can be determined from the table given above and is based upon the total dissolved solids of the water and the diameter of the tank. Magnesium rods should not be installed in larger tanks which are corroded to a point of mechanical failure. If a tank has not reached this point of deterioration it may be possible to weld any leaks and place a cement or preferably a ceramic lining ("elastic porcelain") inside the tank after the steel has been sand blasted to a grey-white metal finish.

Figure 8

Magnesium Rod Installation for Large Hot Water Tanks.


SUPPORT MAGNESIUM RODS WITH RUBBER OR PLASTIC LINED HALF SECTIONS OF 2" DIA PIPE 3" LONG WELDED TO TANK
3-CORED MAGNESIUM RODS
A
A
B
SEE DETAIL "A" FOR TYPICAL CONNECTIONS
SECTION A-A
END B
1¼" EXTRA HEAVY COUPLING
1" EXTRA HEAVY COUPLING
CORED MAGNESIUM ROD
1" PLUG
TANK HEAD
NOTE: PIPE FITTINGS DETAIL A FOR 1.315" ROD PIPE FITTINGS
DETAIL A

d. *Protective Coatings and Linings.* Protective coatings and linings serve as barriers between the material to be protected and the corrosive environment. In many cases, it may be more economical to use a protective coating than to use a more resistant material. Depending on circumstances, protection may be needed either internally or externally. Protective coatings and linings may either be nonmetallic or metallic. Nonmetallic protection includes paint and enamel coatings, cement, masonry, and plastic linings, rubber sheeting, and glass and ceramic coatings. When installing pipe, extreme care must be taken not to scrape off paint or enamel coatings as corrosion will be exceptionally severe where the coating has been damaged. A general rule cannot be made as to the selection of protection methods because each has its application to special circumstances.

(1) *Paints.* Painting or coating is by far the most generally accepted method of corrosion control. Paints must be selected to suit the peculiar conditions encountered. Any foreign material between the metal and protective paint interferes with effective protection. It is therefore important that the surface be clean and free from all loose scale and rust before it is painted. Grease and oil should be removed with a solvent. Oxidized surfaces can be cleaned by pickling, sand blasting, or scratch brushing. A brief description of some of the more common types of coating materials follows:

(a) *Red Lead.* Red lead contains varying percentages of lead oxides (litharge and red lead in a suspension of linseed oil). It has the valuable property of setting up a hard elastic film that excludes moisture and gases likely to cause corrosion. The density of this film depends on the amount of pigment in the paint.

(b) *Iron Oxide.* Red iron oxide is widely used as pigment in the manufacture of protective coatings. There are many grades in varying concentrations of pigment, containing zinc oxide, zinc chromate, or red lead. Such paints, with excellent corrosion-inhibiting values, are used not only as primers but as surface coats over red lead primers.

(c) *Carbon Surface.* Black paints are usually used as finishing coats on steel structures. Carbon pigments, such as carbon black, oil black, and natural and artificial graphite, are usually used as base pigments in these paints. A very tough elastic surface film may be obtained by the addition of litharge and red lead, which give excellent protection over the proper primers.

(d) *Zinc.* Zinc oxide is a rust-inhibiting pigment because iron is cathodic to zinc and will be protected at the expense of the zinc in contact with it. Zinc dust, in specially prepared vehicles, acts to some extent as galvanizing; in fact, a so-called liquid galvanizing product has been prepared from a finely divided zinc dust emulsion in boiled linseed oil.

(e) *Aluminum.* Aluminum paints, which rank high as metal protectors, can be used as finishing coats or as rust-inhibiting primers. They consist of aluminum-bronze powder in a suitable oil-varnish vehicle. They are widely used for coating surfaces submerged in water, with the aluminum powder dispersed in a hard drying synthetic resin vehicle.

(f) *Bituminous.* Bituminous paints are prepared by blending refined coal tar pitch, asphalt, and linseed oil, varnished and thinned down in suitable mineral thinners. They are used in protecting the underground elements of systems from soil and bacterial corrosion. Special blends are used for the protection of water tanks and submerged metal surfaces. American Water Works Association specifications for cast iron water pipe and fittings cover the approved practices for coating cast iron pipe with hot coal-tar pitch varnish. The same association also has approved specifications for coal-tar enamel of steel water mains.

(2) *Wrappers.* Wrappers, a specialized form of protective materials, are used mainly on buried pipelines coated with bituminous compounds. They act as additional protection and are resistant to rot and decay, and to acid and alkaline soils. The asbestos-felt wrappers are flexible and sufficiently sturdy to protect pipeline coating from earth loads and soil stresses. These advantages enable the pipe coating to function properly against corrosion. Severe localized corrosion may occur in areas where the protective coating has been accidentally removed by improper handling during installation.

(3) *Metallic Coatings.* Zinc and lead coatings are the most widely used in waterworks. The use of galvanized pipe (zinc coating) is very common for interior piping and service connections. Galvanizing is particularly applicable to steel and wrought iron and provides good protection against such atmospheric corrosion as is encountered in industrial areas or the sea coast. It is a poor type of protection against agressive and corrosive waters. Other metallic linings and coatings available for protecting piping against atmospheric and special corrosive environments are aluminum, copper, chromium, and nickel. They are applied by spraying, dipping, or electroplating. Certain other processes provide for adding an inhibitor (passivators) to a metal to change the electrochemical potential to a more cathodic value. Examples of this are the Parkerizing, Bonderizing, and Anodizing processes. These alter the surface of the steel and produce a corrosion resistant coating of phosphates or iron oxides on the surface.

(4) *Cement Linings and Coatings.* Cement lined cast iron and steel piping and fittings are commercially available for corrosion protection. Mains and hot water storage tanks may be lined in place by various patented commercial processes at considerable savings over replacement of the facilities. Glass lined hot water storage tanks are readily available on the market and their use is recommended where corrosive water conditions are encountered. Care must be taken in their handling and installation so as not to crack the lining.

e. *Corrosion Inhibitors.* Water is frequently treated to eliminate the causes or reduce the intensity of corrosion. Water may be mechanically deaerated to remove corrosive gases such as oxygen and carbon dioxide or treated chemically to remove these gases. Chemical treatment adjusts the pH of the water and affects the deposition of a solid or gelatinous film on the metal surface. This film, when evenly formed, effectively isolates the metal surface and prevents direct contact between water and metal.

(1) *Anodic Inhibitors.* Soluble phosphates and silicates form films on anodic areas and are known as anodic inhibitors. They prevent corrosion without affecting the potability of water; other compounds such as chromates are toxic and cannot be used in potable waters. Anodic inhibitor dosages must be large enough to stop anodic reaction completely. If the inhibitor does not cover the entire anodic area, the anodic reaction is concentrated and pitting may be greater than when no inhibitor is used. Correct dosage can be found only by trial and inspection.

(a) *Phosphates.* Glassy phosphates such as Calgon, Nalco 918 and similar compounds have wide application in threshold treatment for the prevention of corrosion in both cold and hot water systems. Corrosion inhibition is caused by the adsorption of the phosphate or one of its complexes on the metal surface, forming a protective film. The rate of supplying the phosphate is the important factor rather than its concentration in the water. For initial installations it is desirable to have a high feed (10 ppm) to form a film as quickly as possible. As soon as the protective film has formed the dosage may be reduced to 2 to 4 ppm, depending on what is required to maintain a constant film on the metal. Phosphates are useful in a pH range of 5.0 to about 8.0, with best results obtained when the pH ranges from 6.5 to 7.0. In beginning treatment in corroded pipe systems, difficulty is often experienced with the sloughing off of corrosion products due to the phosphate treatment. Therefore care is required to avoid complaints. Chemical feed is usually by solution feeders identical to hypochlorinators. Since polyphosphate solutions support bacterial growth, use a small amount (2-3 ppm residual chlorine) of hypochlorite in the solution to avoid contamination.

(b) *Sodium Silicate.* The addition of sufficient quantities of soluble silicates to water will produce a precipitation of calcium silicate or other metal silicates on the metal. This method works best where solid corrosion products are present. To maintain a film, sodium silicate must be fed continuously at rates varying from 1 to 30 ppm, depending on requirements. Silicates are best adapted to the treatment of very soft waters. Adding lime to the water raises the pH and helps the formation of calcium silicate coating. The use of silicates in high magnesium waters is not recommended because of the danger of silicate scale formation.

(2) *Cathodic Inhibitors.* Cathodic inhibitors such as calcium carbonate or magnesium hydroxide deposit film on cathodic areas because of the greater alkalinity at these areas. The film excludes

oxygen from cathodic areas, thus reducing the cathodic reaction. The object is to raise the pH of the water above its calcium carbonate saturation value so that a thin calcium carbonate film will be continuously deposited on the metal. Treatment varies with the condition of the water. Hard alkaline waters may require the addition of only slight quantities of soda ash. Soft waters of low alkalinity may need both lime and soda ash. Two main factors must be considered when treating the water, pH and temperature. Slight changes in either could change the water from a corrosive water to a scale-forming one. In conjunction with this treatment, see "Scale". Naturally hard water containing calcium carbonates or silicates often deposits these mineral coatings on metals, thus affording a natural protection against corrosion.

f. *Classification Table*. Table 2 shows a general category classification of waters with respect to corrosion and scaling properties. Suggested inhibitor treatments for corrosion control and threshold treatments for scale control are indicated.

Table 2. CATEGORY CLASSIFICATION OF POTABLE WATERS

Category	Calcium (Ca) ppm	Sulfate SO_4 ppm	Silica SiO_2 ppm	Dissolved Oxygen ppm	Character	Suggested Treatment
1A	0–18	As found	0–15	1–10	Extremely corrosive hot and cold.	*(FOR CATEGORIES 1A-2B).*
1B	0–18	0-25	0–15	0–1	Moderately corrosive hot and cold. Extreme corrosion with CO_2 above 8.0 ppm.	(1) Polyphos. 5–10 ppm (2) Lime to pH 8.5 plus silicates 10 ppm as SiO_2.
1C	0–18	0-60	Above 15	1–5	Slight corrosion on cold; considerable on hot. Aggressiveness reduced, and may not be troublesome due to high natural SiO_2.	(3) Tetra sodium pyrophosphate 8–10 ppm. (4) Di-sodium phosphate anhydrous 5–8 ppm plus soda ash for pH. (5) Lime alone on basis

Description

Scale, as used here, refers to the white crusty matter formed by the crystallizing of dissolved minerals found in natural and lime-softened water supplies. Types of scale include pure mineral scales such as calcium carbonate and calcium sulfate, mixed scale and corrosion products, and combinations of minerals cemented together by a binder such as silica. Bacteria and algae contribute to the scale problem by forming growths in pipes and further restricting flow. Some substances, such as silica, produce a hard "baked-on" type of scale; others produce a relatively loose, sludge-like deposit.

Scale Removal

Scale deposits can be removed by mechanical cleaning if their adherence properties are not great. Reversing flow may remove loose scale. Calcium sulfate scale can be cracked off heating coils by a rapid change in temperature. Thin calcium sulfate scale can be removed by treatment with one of the orthophosphates. Carbonate scale can be removed by the application of commercial grade (muriatic) hydrochloric acid containing a good inhibitor. Inhibitors are polar molecules which permits them to be attracted to the iron (negative metal charges) of the boiler and establish a thin protective film. Two general methods used in acid cleaning of boilers are the "circulation" and the "fill and soak". Generally, it is preferable to have this work performed by outside contract.

CORROSION AND SCALE CONTROL ACTIVITIES AND TESTS

Corrosion and Scale Control Activities

a. *Organized Studies.* Continuing studies of corrosion and scale problems are made by Government agencies such as the Bureau of Standards and the National Research Council, by industry, and by such technical groups as the National Association of Corrosion Engineers, the Electrochemical Society, and the American Society for Testing Materials.

b. *Water-Formed Deposits Analyses.* In initiating corrosion and scale control programs, it is often necessary that the character of the corrosion or scale products be known. This is particularly true when it is desired to initiate specific treatment or to clean and remove built-up products through the use of chemical cleaners and inhibitors. To assist installations and commands in solving corrosion and scale problems, laboratory services for analyzing water-formed deposits for chemical content are available from the U.S. Geological Survey, Quality of Water Branch, Washington 25, D. C. The following instructions will be followed in the collecting and submission of samples to the U.S. Geological Survey for analyses:

(1) Select a sample from an area representative of the trouble experienced; usually an area of failure or heaviest deposition of scale or corrosion products. When the affected area is not available for spot sampling, select a sample from an adjacent area that is directly related to the area of failure or the area of principle concern and that will at the same time provide an adequate sample.

(2) Deposits should be carefully removed and maintained in their original state as much as possible. Contamination with any outside material should be avoided. A quantity amounting to 1 ounce to 3 ounces is usually sufficient. It is desirable to submit a small section of the affected area of pipe containing the deposit.

(3) Containers for shipping samples should be wide mouthed glass bottles or plastic vials. Metal containers can be used if the samples are dry when collected. Special containers and tags are available from the U.S. Geological Survey on request.

(4) A description of the sample should be attached to the container.

c. *Command Base Activities.* Installations in different locations encounter corrosion problems of varying kinds and scope. Technical problems of a complex nature should be referred to command level; command engineers are available for technical assistance on request by an installation.

Testing for Corrosion and Scale

Many corrosion and scale problems appear only in the course of operating a utility or other system. Causes must be accurately determined before effective remedial action can be taken.

a. *Laboratory.* Of the two general classes of corrosion and scale tests, laboratory and field, the field test is more accurate and dependable because actual service conditions are involved. It is essential that laboratory tests simulate these service conditions as closely as possible and that no new factors be added; for example, continuous service should be used in a laboratory test to imitate the long-term effects of corrosion and scale on equipment in normal service.

b. *Field.* Field tests for corrosion and scale may be made as follows:

(1) *Testers:*

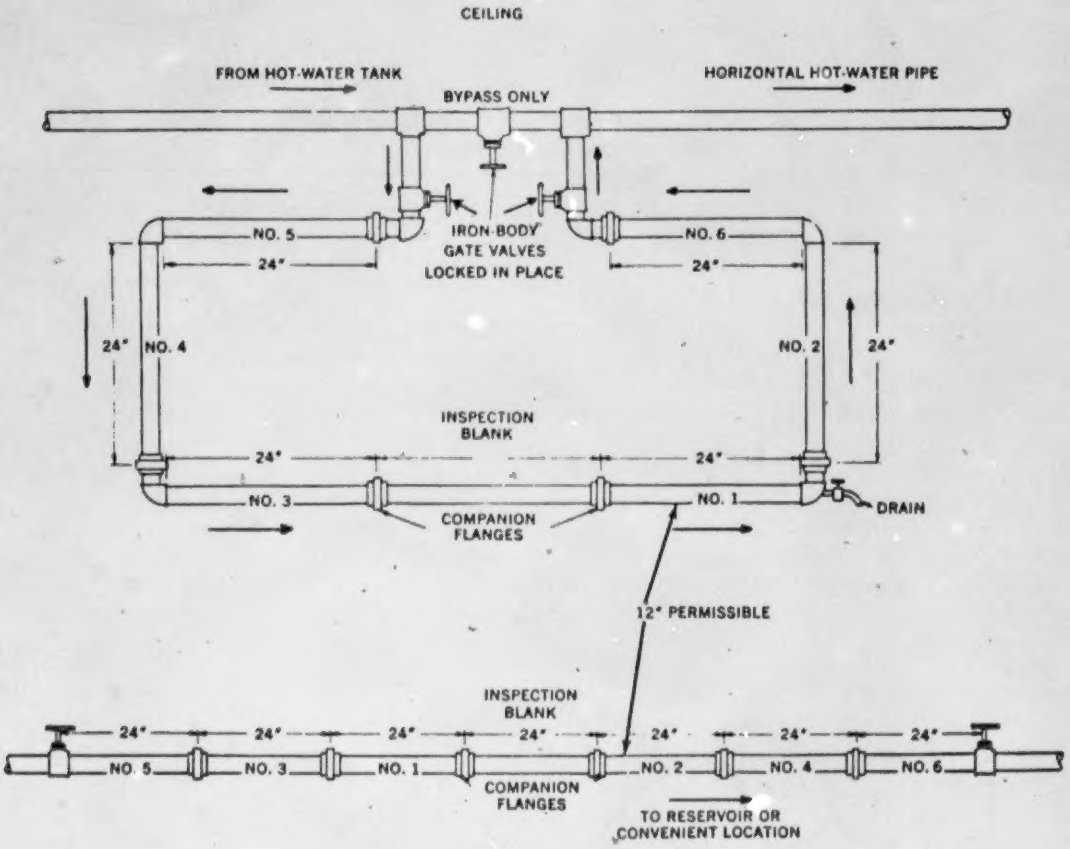
(a) *Nipples.* Test or observation stations should be established at each installation. Stations should be located in the hot-water heating plant of a dining facility near the point where chemicals are added to the water supply, and about the middle and the far end of the installation distribution system. At least one set of cold-water nipples should be installed; if water is treated for corrosion control, a second set should be installed after chemical treatment. Figure 9 shows the layout for test nipples used to check on corrosion. Pipe diameters are the same as that of the discharge pipe from the hot-water tank, so velocity of flow will remain the same. Test nipples should be the same as those existing, or should be installed with special insulated connectors. Installations having nonferrous piping need not install test sections.

(b) *Strips.* Test strips may be installed at selected locations if local conditions warrant.

(c) *Evaluation.* The degree of scale control is shown by the thickness of scale in the test nipple. Corrosion and scale deposits are cleaned off the strips and nipples and the depth of pitting is measured to determine the corrosiveness of the water. If a number

Figure 9

Layout of Test-Pipe Loops to Observe Treatment of Corrosive Water.


CEILING
FROM HOT-WATER TANK
HORIZONTAL HOT-WATER PIPE
BYPASS ONLY
IRON BODY
GATE VALVES
LOCKED IN PLACE
NO. 5
NO. 6
24"
24"
24"
NO. 4
NO. 2
24"
INSPECTION
BLANK
24"
24"
NO. 3
NO. 1
DRAIN
COMPANION
FLANGES
12" PERMISSIBLE
INSPECTION
BLANK
24"
24"
24"
24"
24"
24"
24"
NO. 5
NO. 3
NO. 1
NO. 2
NO. 4
NO. 6
COMPANION
FLANGES
TO RESERVOIR OR
CONVENIENT LOCATION

of test nipples of the same material are installed in a test loop when treatment is started, it is possible to obtain progressive data on the results. For example, the depth of pits may decrease and the pits widen as treatment progresses. Each subsequent test nipple removed may show this.

(2) *Steel Coil Test.* Steel coils can also be used to test the effectiveness of corrosion and scale treatments.

(3) *Steel Distribution System Test Points.* At installations with steel pipe distribution systems, a bypass may be built at any convenient point. A section of the bypass should be removed periodically and examined to determine whether any change in corrosion control is necessary.

(4) *Samples of Test Specimens.* Samples of test specimens from building plumbing and distribution system should be retained at the installation as a permanent record of scale and corrosion conditions.

CORROSION AND CONTROL*

Corrosion, the process of slow oxidation of metals, is not only a malignant destroyer in our water works industry, but it creates other objections which may even be more serious. Consumers' complaints are caused by iron and manganese, offensive particulate matter, and corrosion products in drinking water. There are also complaints of corrosion-caused rapid deterioration of plumbing. The loss of metal pipe of the water system may not be nearly so damaging as the loss of carrying-capacity caused by pipe tuberculation.

Because iron is the most widely used of the many metals utilized by the water works industry and because corrosion differs with these metals, this discussion will be primarily about corrosion of iron. We know that iron exists naturally as an ore, e.g., Hematite (Fe_2O_3), Magnatite (Fe_3O_4), Iron Pyrites (FeS_2) and Siderite ($FeCO_3$). Iron pipe is mostly elemental iron (Fe^o). This iron, in the uncombined state, is unstable and tends to oxidize, or rust. It tends to return to the more stable state such as one of the ores mentioned above.

Theory of Corrosion

Much has been written about corrosion in the last few decades. Many theories regarding its cause and how it might be controlled have been expounded. The electro-chemical theory of corrosion is perhaps the most commonly accepted. To simplify, this theory might be described in four steps:

1. $Fe^o \longrightarrow Fe^{++} + 2(e)$

Iron goes into solution as an ion (Fe^{++}) by giving up 2 electrons. A negative charge tends to build up and retard the reaction but hydrogen ions (H^+) in solution will take up these electrons and form hydrogen gas.

2. $2H^+ + 2(e) \longrightarrow H_2$ (gas)

The hydrogen gas coats the metalic surface and separates it from the water. This action, called polarization, tends to retard corrosion, but oxygen dissolved in the water reacts with the gas and depolarizes the metalic surface.

3. $2H_2 + O_2 \longrightarrow 2H_2O$

In our case of corrosion of water pipe the depolarizing agent is one of the principal factors determining rate of corrosion.

*Source: Water Works Operators Manual, Alabama Department of Public Health, Bureau of Environmental Health, Third Edition, 1972, Chapter XX.

4. $2Fe^{++} + \frac{1}{2}O_2 + 4(OH)^- \longrightarrow 2H_2O + Fe_2O_3$ (rust)

OR

$2Fe^{++} + \frac{1}{2}O_2 + 4(HCO_3) \longrightarrow 2H_2O + 4CO_2 + Fe_2O_3$
Fe_2O_3 (rust)

OR

Iron may react with silicates, phosphates or other ions in the water.

Factors Influencing Corrosion In Water Pipes

Dissolved Oxygen---This is the most important factor in water pipe corrosion. Removal of oxygen is practical for corrosion control in some special cases but such a procedure on a municipal scale is not feasible. Water devoid of oxygen is usually non-corrosive when the pH is higher than 7.0 but unless such water is kept away from contact with air it is practically impossible to prevent some oxygen absorption. There are other oxidizing agents which can substitute for oxygen and influence corrosion. Such agents include nitrates and chlorine.

Carbon Dioxide---Although carbon dioxide reacts with water to form a weak acid which enters the corrosion reaction, its most important role is the influence it has on the products of corrosion. As will be discussed later, carbon dioxide enters into the equilibrium of alkalinity and calcium hardness and thereby influences the rate of corrosion.

Alkalinity, Calcium, and pH---These factors influence the formation and maintenance of a calcium carbonate scale and thereby affect the rate of corrosion. In the range encountered in water supply, pH alone does not directly affect the rate of corrosion but at pH values below 4 for iron and 7.0 for copper the rate of corrosion is increased at a logarithmic rate with a decrease in pH. On the other hand pH does influence the solubility of calcium carbonate, viz., raising pH lowers solubility, and consequently the calcium carbonate scale can be created to influence corrosion.

Rusting will soon commence when water is turned into a pipe, but if the water is supersaturated with calcium carbonate a precipitate will form and mix with the corrosion products. The resultant scale will retard corrosion and continued precipitation of calcium carbonate will create a scale to cover the surface of the pipe. If the scaling continues, the entire pipe can be completely filled. Reduced cross-sectional area can be just as damaging as corrosion, consequently these factors of pH, alkalinity, and calcium must be kept in balance in order that both corrosion and scale formation will be minimized.

Temperature---This factor causes several conflicting effects. Increased temperature normally causes increased corrosion. The capacity of water to absorb oxygen is reduced with increased temperature, but in the ranges encountered in municipal water supplies the former factor is predominant. As temperature increases solubility of calcium carbonate is reduced and consequently rate of scaling increases. This will be particularly noticeable in hot water heaters when the water is high in calcium hardness and alkalinity.

Velocity of Flow---An increase in the flow of water through the pipe accelerates corrosion by increasing contact of oxygen with the pipe surface. This increased contact results from both the increased amount of oxygen available (more water) and as a result of turbulence an increased amount of water contacting the pipe surface. Turbulence also will adversely affect the creation of a protective film on the pipe surface.

Nature of the Metal---Rate of corrosion varies according to the kind of metal and even with different types of the same metal. Among the irons, or ferrous metals, wrought iron is generally more resistant to corrosion than cast iron and cast iron is more resistant than ordinary steel. Of course, stainless steel is noted for its high resistance to corrosion, but because of high cost, it is used only in special cases.

Bacterial Corrosion---There is much yet to be learned about the relationship of bacteria and corrosion. In special situations bacteria can play a very important role in causing corrosion; e.g., some sulphur bacteria can produce acid from sulphides or sulphates in water. Such acids created may reduce the pH sufficiently to cause severe local corrosion. Normally chlorination can prevent such a development but if growth of these bacteria gets started corrosion products can quickly isolate the growth and build a protection against disinfection. When this occurs, an alkali is added to adjust the pH above 8.2.

Diagnosing Corrosion

When you are bombarded with complaints that red particles are coming from consumers' water faucets and you know there is no iron in the raw water you can assume that corrosion is taking place in your distribution system or in the consumers' plumbing. If such red water comes from the hot water taps only, the corrosion is not nearly so severe as when it also comes from the cold water taps. You cannot assume, however, that you are free from corrosion if you do not receive these complaints. When the pH of the water is high, corrosion may be forming tubercules and as velocity of flow is low, rust may be remaining in the system. If this is the case, flushing mains will reveal the corrosion.

When pipe is removed from the distribution system it should be inspected to reveal the condition of corrosion or scaling.

In many water systems hydraulic surveys are made to determine carrying capacity of water mains. By such surveys reductions in carrying capacity as created by corrosion or scale formation can be detected and the results of corrective action can be evaluated.

Another scheme for checking on corrosion is to determine the reduction in dissolved oxygen of the water as it progresses through the distribution system. If oxygen is being removed there is a good chance it is going into the corrosion reaction.

Methods of Water Pipe Corrosion Control

Controlling corrosion in any particular water system must of necessity be based on evaluation and control of the factors influencing that corrosion. Some of the factors are not amenable to control but their influence might be eliminated by special means. There is not much that can be done about dissolved oxygen but the use of pipe materials other than metal can eliminate the corrosion. Such a course may lead to problems other than corrosion and will not eliminate corrosion of consumers' plumbing. There is not much that can be done about temperature, but its effect is not too important in the distribution system. A well treated water which is continuously and adequately disinfected (with a chlorine residual maintained throughout the system) should be free of the bacterial corrosion. Maintaining a system adequately sized for the demand placed upon it will prevent troubles from high flow velocities.

The major activities for water pipe corrosion control could be classed under two headings: Choice of Pipe and Water Treatment. The most important of these two is Water Treatment.

Choice of Pipe---In choosing pipe, as pointed out above, other problems than corrosion control may be encountered when other than metal pipe is selected. An evaluation of involved problems will indicate which materials to use for any particular job.

The pipe corrosion problem can be eliminated by separating the pipe from the water. This can be done by providing the pipe with a protective lining. New pipe now being laid is lined pipe and considerable pipe requiring cleaning is being lined in place. There is a wide variety of linings available and it may be desirable to test proposed linings with the water which the pipe is to convey before a selection is made. Pipe lining materials include products of portland cement and bituminous or asphaltic compounds.

Water Treatment---Treatment of water for corrosion control is most satisfactorily accomplished when the water is maintaining in chemical balance. By this means most of the corrosion problems can be solved. A treatment method short of this chemical balance but found effective is the use of specific inhibitors for corrosion control.

Use of Inhibitors---Numerous inhibitors have been developed to make waters non-corrosive; e.g., the use of inhibitors to protect automobile radiators. Most of these materials have specific application, and, because of toxic effects, are not suitable for use in municipal supplies. Some inhibitors, however, are being used satisfactorily in municipal water supplies; two of these are sodium-silicate and glassy phosphates.

The continuous feed of soluble silica provides corrosion control by means of a thin film, claimed to be iron-silicate, on the surface of the pipe. On commencing silica feed, a concentration as high as 8 ppm is used to rapidly establish the film throughout the distribution system. When this film has been established the feed rate is reduced to approximately 2 to 4 ppm.

Glassy phosphates have proved suitable to control corrosion under some conditions. Control by this method also depends upon maintaining a protective film on the pipe surface. A high initial feed is used to provide a rapid establishment of a film throughout the system. This film is maintained by a continuous feed of from 1 to 2 ppm. Inasmuch as the maintenance of this film depends on a continuous application of freshly treated water a program of flushing dead ends of the system must be practiced or control will be interrupted. In order for the glassy phosphates (sodium hexamethaphosphate) to be effective, the pH of the water must be maintained as high as 6.8 to 7.2. Experience in Alabama indicates that the problem may become more acute, rather than better, if the hexametaphosphate is added to water with a pH lower than 6.5.

Glassy phosphates are used for preventing scale formation. Even low concentrations inhibit crystal formation and as a result any scale which might form will be of much finer particles, more dense and more evenly distributed throughout the distribution system.

A special application of this treatment is found in household hot water systems where temperature effect becomes important. By treatment with glassy phosphates, both corrosion and scaling might be controlled. Here the dosages are much higher than those used in municipal water systems.

Treatment to Maintain Chemical Balance---Before going into treatment to produce a chemically balanced water or one which will provide a light carbonate protective coating, we should discuss some of natural water's chemical characteristics.

First, let us take "pH". We have been mentioning pH, but just what is it? Actually, it is a measure of the hydrogen ions in solution. By definition it is the negative logarithym of the hydrogen ion concentration. The value of pH does not tell the magnitude of the alkalinity or acidity of water but it gives a measure of the water's acid or alkaline reaction. Actually, a water can, and most always does, have an alkalinity and an acidity at the same time. The pH is measured either by a pH meter or by a visual colorimetric test.

Alkalinity in water is the amount of titrable base contained (i.e., how much base there is to react with acid). The alkalinity is determined by adding known quantities of a standard acid to a given volume of water until the alkalinity is neutralized to a given end-point, usually determined by a colorimeter indicator such as Methyl orange. The alkalinity is calculated from results of this neutralization (or titration) and reported as ppm $CaCO_3$. Alkalinity usually consists of bicarbonates (HCO_3), carbonates (CO_3), or hydroxides (OH), and less frequently as borates, silicates, and phosphates.

Acidity of water is the titratable acid and is determined by the amount of base required to neutralize this acid, or raise the pH to a given end-point. In natural waters this acidity is generally caused by carbon dioxide in solution (carbonic acid). Mineral acids, getting into the water as pollution, of course are included in the total acidity.

A relationship of the constituents making up the acidity and alkalinity in water may be shown thusly:

$$\begin{tikzcd}
H_2O \arrow[d, shift right] & + & CO_2 & \\
CaCO_3 \arrow[u, shift right] & + & H_2CO_3 \arrow[r, shift left] & Ca(HCO_3)_2 \arrow[l, shift left]
\end{tikzcd}$$

This equation shows an equilibrium condition wherein the amounts of carbonate, carbon dioxide, and calcium held in solution are interdependent. If water is deficient in alkalinity or, in other words, there is an excess of carbon dioxide, there will be a tendency to dissolve calcium carbonate (or iron, or anything available) to neutralize this excess carbon dioxide. If the water has some alkalinity and calcium but is deficient in carbon dioxide it can be stabilized by the removal of this excess carbon dioxide. If more than the excess carbon dioxide is removed, e.g., by aeration, boiling, or addition of lime, the water will tend to precipitate calcium carbonate. When water in the distribution system is on the side of the balance tending toward precipitation, a carbonate scale will be formed on the water pipe. When the water is on the side tending toward solution of calcium carbonate the scale will be dissolved, or, in the absence of scale, the iron pipe will be corroded.

There are formulations showing the relationship of calcium carbonate alkalinity and carbon dioxide. Such a relationship can also be shown in the form of alkalinity vs. pH. By using a curve showing this latter relationship a rough estimate of a given water's corroding or scaling characteristics can be made, even when only pH and alkalinity are known.

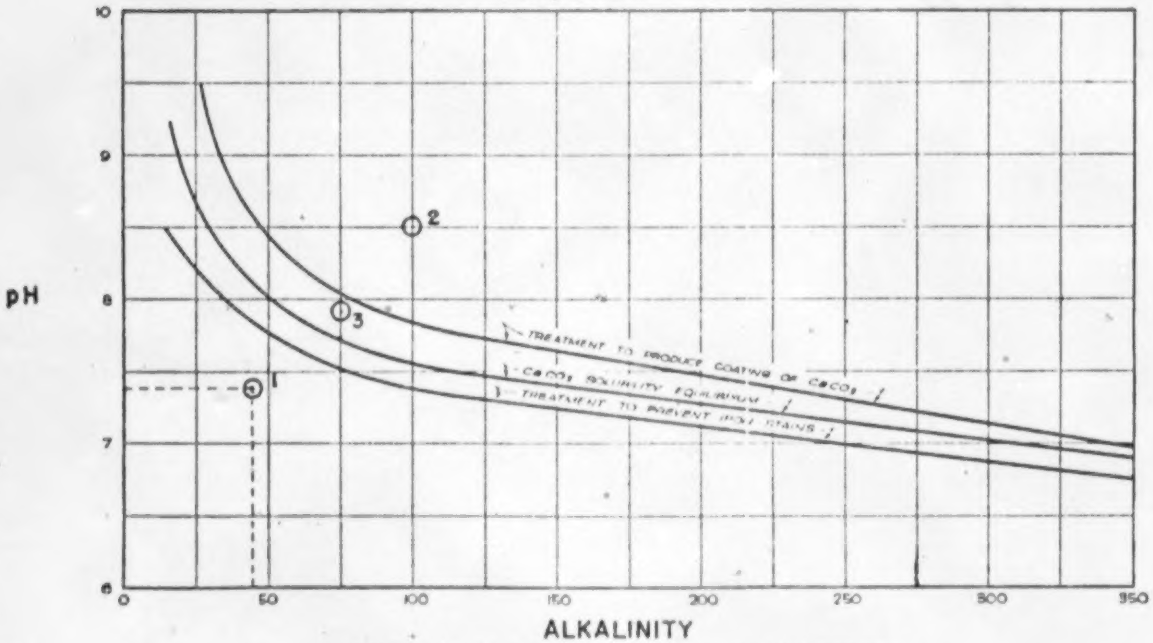
Inasmuch as a number of other factors are also involved a better appraisal of the water's corroding or scaling characteristics can be made by use of the "marble" or "calcium carbonate stability" test.

Another method of appraising a water's corroding or scaling tendencies is by means of Langelier's pH of calcium carbonate saturation (pH_s). This value may be calculated when the following are known: pH, total dissolved solids, water temperature, calcium hardness (as $CaCO_3$), and Methyl orange alkalinity (as $CaCO_3$).

Figure 1 is a chart showing the relation of pH and alkalinity to the proper treatment of water to prevent corrosion. This, if used by the operator, will serve as a guide to him in the proper addition of post alkali, for corrective treatment. The reason for making the total alkalinity and pH tests on the finished water is to obtain data which will possibly tell the operator whether or not the water he is sending out is likely to corrode the distribution system and cause "red" water or whether the water contains so much alkali that it will build up on the lines and cause stoppages. The other tests such as phenolphthalein alkalinity, carbonate stability and carbon dioxide give information which, if used along with the total alkalinity and pH values, will enable him to further refine his control over the water.

The chart (Figure 1) is used in the following manner: Suppose the operator makes total alkalinity test on finished water and finds the result to be 45 ppm and the pH value to be 7.4. Using the chart he locates 45 on bottom line labeled "alkalinity" and 7.4 on the vertical line labeled "pH". The intersection of two lines, one drawn upward from 45 and the other extending to the right from 7.4 will be at point located at (1). Since this line falls below the curve labeled "Treatment to prevent iron stains", the water will more than likely be corrosive and "red" water complaints can be expected if the treatment continues in this manner.

Figure 1
Baylis Curve



RELATION OF pH AND ALKALINITY TO THE PROPER TREATMENT OF WATER TO PREVENT CORROSION
The calcium carbonate solubility curve shown is based upon calcium carbonate in distilled water at 22°C
It is approximately correct for natural waters low in soluble salts but should be checked when used on waters containing considerable dissolved compounds.

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Now suppose the operator changed his lime feeder to add considerably more lime. After sufficient time for the change to affect the finished water he makes his tests again. This time suppose his results were: Total alkalinity 100 ppm and pH 8.5. Entering the chart at the proper places lines from these points alkalinity and pH scales would intersect at (2). Since this point is above the curve "Treatment to produce coating of $CaCO_3$" too much lime is being added and may possibly deposit in the distribution system lines. The lime feed should be reduced.

The operator should continue to change the lime feeder until the two points meet on the curve labeled "$CaCO_3$ solubility equilibrium" or between it and the curve above. For example suppose he found after further changes in the feeder and tests the results were: Total alkalinity 75 ppm and pH 7.9. Entering the chart this time, the point of intersection would be at (3), indicating that the water was being treated as it should be. With the carbonate stability or marble test he could determine exactly how much lime should be added.

LABORATORY EQUIPMENT AND SUPPLY REQUIREMENTS FOR A SMALL WATER TREATMENT PLANT PROVIDING COMPLETE TREATMENT

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Drinking Water Program
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The following laboratory equipment and supplies are required of small water treatment plants for determining the physical-chemical parameters of chlorine residual (free and total), turbidity, pH, alkalinity, free carbon dioxide, hardness, fluoride ion, and performing the "jar test" for proper coagulant dosage.

Vendors not noted below, but capable of supplying equipment equal to that which is listed are acceptable. However, the Environmental Health Services Laboratory (telephone 348-0197) should be consulted before purchasing agreements are made with these vendors to confirm that equivalent equipment is being offered.

Item Number	Quantity	Items
1	1	Copy of latest edition of "Standard Methods for the Examination of Water and Waste Water", American Public Health Association, Inc., 1015 Eighteenth Street, N.W., Washington, D. C. 20036
2	1	Balance, Harvard Trip Scale, single beam, Ohaus #1450
3	1 set	Technical weights, metric, class C, 1 gm to 1000 gms
4	1	pH meter - accuracy ± 0.05 unit, scale readability ± 0.1 pH unit. Instruments meeting specifications of the Corning model 7, the Orion model 601, Beckman Zeromatic, Fisher Accumet, or equivalent are satisfactory. Laboratories purchasing a new pH meter are strongly advised to purchase one capable of functioning with specific ion electrodes.
5	3	Burettes, Automatic, with stopcock, for dispensing sulfuric acid and hardness reagent, also N/44 sodium hydroxide, 50 ml. capacity, 1/10 ml. subdivisions, complete with rubber bulb and reservoir.

LABORATORY EQUIPMENT AND SUPPLY REQUIREMENTS FOR A SMALL WATER TREATMENT PLANT PROVIDING COMPLETE TREATMENT

Item Number	Quantity	Items
6	1	Burette Support, with large white porcelain base and a 3 way castaloy burette holder for assembling and steadying automatic burettes with reagent reservoirs.
7	1	Therometer, - 10° to 110° C.
8	4	Casseroles, porcelain (Coors #180), 210 ml.
9	6	Stirring rods, glass, 3mm diameter
10	6	Pipettes, water analysis, 10 ml. capacity
11	2	Cylinders, graduate, 100 ml. capacity, 1.0 ml. subdivisions
12	3	Flasks, Erlenmeyer, wide mouth, 250 ml.
13	6	Beakers, Griffin, "Pyrex", 1000 ml. (for jar test)
14	6	Bottles, dropping T. K., 60 ml.
15	1	DPD Chlorine residual colorimetric test kit. LaMotte, Hach, Taylor or equivalent are satisfactory kits. Kits must include free and total residual capabilities.
16	1	Turbidimeter (Nephelometric only) - Hach Turbidimeter 2100 A, Sargent-Welch, Turner Designs, Monitek, or equivalent. TURBIDITY ANALYSIS IS REQUIRED ONLY FOR THOSE SUPPLIES TREATING SURFACE SOURCE WATERS. Calibration of sealed liquid turbidity standards - sealed liquid turbidity standards purchased from the instrument manufacturer must be checked frequently for settled and floating particles in order to monitor their eventual deterioration. These standards should be replaced when any major change from the previous calibration occurs or sediment appears. Solid turbidity standards composed of plastic, glass, or other materials are not reliable and should not be used.

LABORATORY EQUIPMENT AND SUPPLY REQUIREMENTS FOR A SMALL WATER TREATMENT PLANT PROVIDING COMPLETE TREATMENT

Item Number	Quantity	Items
17	1	Distilling apparatus, automatic, bantam still, capacity 1.8 liters per hr., Barnstead or equivalent with 2 gallons of distilled water storage. (Preiser Scientific - Catalog #17-8360 or equal)
18	2	Propipette Bulbs
19	1	Filter photometer, AC, Hach or equivalent (Hach No. 110400, 115 V, 60 Hz) with capabilities for iron, manganese, fluoride (SPADNS) method, Nitrate (cadmium reduction), with necessary reagents.
20	1	Specific ion meter - expanded scale millivolt capability, readable and accurate to 0.1 millivolt. Instruments meeting the specifications of the Orion model 407 A, the Orion model 801 A, the Corning model 110, or equivalent are satisfactory, complete with specific ion electrodes. FLUORIDE ANALYSIS REQUIRED ONLY FOR THOSE SUPPLIES FLUORIDATING.
21	1	Automatic Batch Fluoride still, Hach or equivalent, Hach Catalog #592. NOTE: If the SPADNS colorimetric method for fluoride is used, all samples must be distilled before this determination is made. If the specific ion electrode method is used (as mentioned in item #20 above), this piece of equipment can be omitted.
22	1	Stirring apparatus for "jar test", 6 unit, with variable speed control and light reflector.
23		Chemical Supplies:
	32 oz.	Sulfuric Acid, N/50 solution
	32 oz.	Sodium Hydroxide, N/44 solution
	16 oz.	Methyl Orange indicator solution
	16 oz.	Phenolphthalein indicator solution
	8 oz.	Sodium Arsenite reagent
	16 oz.	Buffer Hardness reagent
	2 vials	Hardness indicator powder (EDTA)
	32 oz.	Hardness Reagent (titrating)
	32 oz.	SPADNS Reagent
	16 oz.	Fluoride solution, standard 1.0 ppm
	1 gal.	TISAB (for use with specific ion electrode

LABORATORY EQUIPMENT AND SUPPLY REQUIREMENTS FOR A SMALL WATER TREATMENT PLANT PROVIDING COMPLETE TREATMENT

Item Number	Quantity	Items
23 cont.		method for fluoride analysis).
	1 pint	pH Buffer 7.0
	1 pint	pH Buffer 4.0

LABORATORY FACILITIES - General Notes

1. Adequate laboratory space should be made available. As a guideline, there should be approximately 150 to 200 square feet per person, regardless of test, and approximately 15 linear feet of usable bench space including a sink with hot and cold water taps.

2. Raw and post-settling water taps should be provided in the laboratory area for quality control monitoring.

3. Sufficient electrical outlets must be provided for operating the laboratory equipment.

4. GLASSWARE - all glassware purchased for laboratory use should be composed of PYREX or KIMAX type glass. This type of glass is more resistant to damage by heat, chemicals, and abuse than is regular soft glass. All volumetric glassware should be marked CLASS A, denoting that it meets federal specifications for volumetric glassware, and need not be calibrated before use.

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